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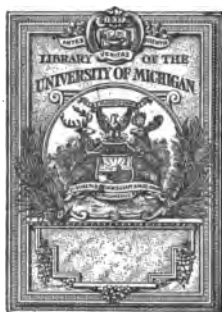
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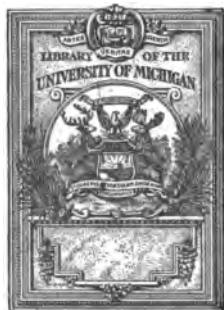
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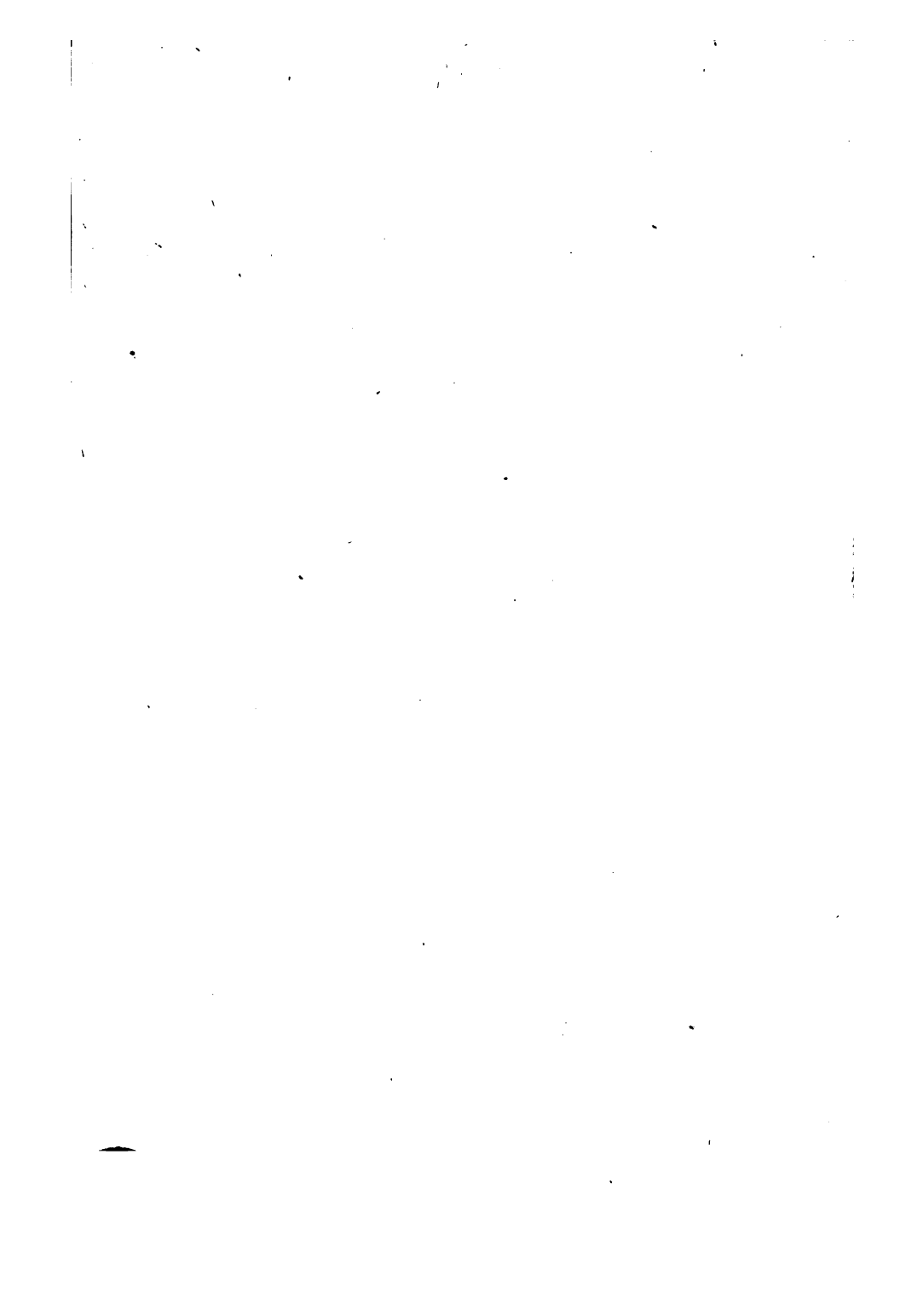
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Compliments of
Henry M. Wheelery
August 5, 1919.

CHEMICAL LECTURE NOTES.

TAKEN FROM PROF. C. O. CURTMAN'S LECTURES
AT THE ST. LOUIS COLLEGE OF
PHARMACY.

BY

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FOURTH EDITION.

WITH OVER ONE HUNDRED ILLUSTRATIONS.

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PREFACE TO FIRST EDITION.

The student who takes careful notes of the lectures, and refers to them repeatedly while studying, is sure to gain a better understanding of the subject than the same person could have done without the aid of notes. When first engaged in conducting the Quiz Class in Chemistry at the St. Louis College of Pharmacy, I took notes of the lectures for the purpose of having the review follow closely after the instruction. In 1884 Professor Curtman entirely rewrote his lecture notes, so as to embody in them a variety of information, gathered from scientific journals and other available sources, not readily accessible to the student, and to form a complete syllabus of the subjects taught. These notes I obtained permission to copy for use with the Quiz Class, and for publication in the columns of the *National Druggist*, of which I had then become editor.

In the department of "Chemical Physics" I have supplied various paragraphs from other sources than these notes, and have prefixed in many places letters and figures, to give a systematic arrangement where the professor had given only an enumeration of subjects by suggestive words, to remind him of facts or illustrative experiments to be brought before the class.

In the chapters on "Metalloids" I have closely followed Professor Curtman's notes, supplying from time to time a sentence to make clearer the meaning of words and formulæ, plain or dissected, which, while sufficient as a memorandum for the teacher, would, without explanation, appear obscure to the student.

The notes have been published in this form more especially for students of pharmaceutical and medical colleges. I believe, however, that they will prove of assistance to all who

desire to study Chemistry or refresh their memory on the subject.

The illustrations are supplied entirely by me, and will recall many class experiments to students, and, together with the various memoranda introduced, furnish interesting information that has been gathered from various sources. Considerable pains have been taken in the pages on Chemistry to display such important subjects as symbol or formula, atomic or molecular weight, etc., immediately following the name of each element, and of the more important compounds. The explanation of the construction of formulæ by means of dissection, which has been carried out in the notes, is a prominent feature of Professor Curtman's lectures.

The index is sufficiently complete to enable the reader to refer to any of the compounds mentioned in the work.

H. M. WHELPLEY.

ST. LOUIS, MO., December 1, 1886.

PREFACE TO SECOND EDITION.

The benefit that students have derived from the first edition of this work, and the demand made by druggists and physicians in general, has induced me to revise the volume, and add Professor Curtman's Notes on the Meta's. This has increased the size of the book about seventy pages, and completes the Notes up to the Lectures on Organic Chemistry.

The index has been thoroughly revised, in conformity with the new matter added.

H. M. WHELPLEY.

St. Louis, Mo., October 1, 1888.

PREFACE TO THIRD EDITION.

The new edition of this work is substantially the same as the preceding one. The duties of a busy life have prevented me from preparing and adding the Notes on Organic Chemistry as originally intended.

The favorable reception given the first and second editions and the demand for more copies has induced me to have the present edition published.

H. M. WHELPLEY.

St. Louis, Mo., April 1, 1890.

PREFACE TO FOURTH EDITION.

The continual demand for this work has induced me to supply a new edition.

I trust the introduction of the Metric System Tables and a List of Latin and Greek Numerals will render the volume more serviceable to the Pharmaceutical and Medical Professions.

H. M. WHELPLEY.

ST. LOUIS, MO., December, 1892.

CHEMICAL LECTURE NOTES.

INTRODUCTION.

I. DIFFICULTIES SUPPOSED TO EXIST IN THE STUDY OF CHEMISTRY.

1. *Misapprehension in regard to the extent of study by medical men or by pharmacists.*
2. *Frequent abuse of the word "practical," as opposed to "theoretical," by men ignorant of their profession, or too indolent to keep up their studies.*

Chemistry is one of the foundation stones upon which both practical medicine and successful pharmacy depend, and its importance is continually growing.

II. THE REAL DIFFICULTIES IN THE WAY OF THE STUDY OF CHEMISTRY.

Wrong methods of study.

Memorizing formulæ from text-books; illustrations that do not illustrate.

Want of perception of practical applications, etc.

Lack of interest by students unless they handle chemicals and apparatus themselves.

Importance of laboratory work, both for acquaintance with material and training in practical methods, and for better understanding of accompanying lectures.

III. PLAN OF SUBDIVISIONS OF LECTURES IN COURSE.

1. *Chemical physics.*
2. *Chemistry of metalloids.*
3. *Chemistry of metals.*
4. *Organic compounds.*

IV. NECESSITY OF STUDY OF CHEMICAL PHYSICS.

1. *Former and present relations of natural philosophy and chemistry.*
2. *Formerly quite distinct departments; line easily drawn; now much common territory.*

3. *Similar experience in other departments of natural history; borderlines being effaced with advancing knowledge; botany and zoology; difficulty of assigning places of bacteria, etc.*

4. *Growing importance of physical and chemical processes in explanation of chemical action.*

5. *Hence, necessity of study of natural philosophy and of teaching especially chemical physics as preparatory to chemistry proper.*

6. *More accurate knowledge of gases in modern chemistry.*

7. *States of aggregation; Kinetic theory; frequent or rarer impulses of molecules the cause of different pressures.*

8. *Air to be studied as the prototype of mechanical properties of gases.*

9. *Water to be taken as an example for some of the mechanical properties of liquids.*

10. *General properties of matter and its relation to forces to be studied before the special actions of affinity can be understood.*

11. *Hence, the course will embrace—*

a. General properties of matter.

b. Gravity—absolute and specific

c. Pneumatics in short outline.

d. Hydrostatics in short outline.

e. Cohesion, adhesion, etc.

f. Heat and light.

g. Electricity and magnetism.

CHEMICAL PHYSICS.

I. GENERAL REMARKS.

1. *Definition of Matter*.—Everything possessing extension and impenetrability. [Webster: Everything visible and tangible.]

a. Example of matter: Nitrogen dioxide in flask not visible or tangible, yet impenetrable, and has extension, shown by inversion of flask under water. By admission of oxygen, red fumes of nitrogen tetroxide are formed.

b. Kinds of matter: Simple and compound.

2. *Definition of Element*.—Homogeneous or simple matter, resisting all efforts to be divided into heterogeneous parts.

a. Number of elements: Seventy-two (1890), to which new additions are occasionally made; some yet undecided.

II. EXPLANATION OF TERMS.

1. *Moles*.—A mass.

2. *Molar*.—Mechanical motion.

3. *Molecule*.—Smallest particle of matter capable of existing in the free state. Physical unit.

4. *Atom*.—Smallest particle of simple matter capable of combining with another.

5. *Ultimate Constitution of Matter*.—Hypothesis of molecular structure; atoms in molecule. Ether envelope.

6. *Physics*.—Changes affecting matter in mass or in its ultimate particles (molecules) resulting from the various physical forces: Motion, attraction, gravitation, electricity, magnetism, light, and heat; also general properties.

7. *Chemistry*.—Special properties of the different kinds of matter; changes produced by affinity in the arrangement of atoms within the molecule.

Distinction between natural philosophy and chemistry gradually blending, and transition of limits.

8. *Relation of Matter and Force*.—Various hypothesis and theories. Force either existing as separable and extraneous to matter, or as residing in it. Separation assumed only for convenience of study.

III. GENERAL PROPERTIES OF MATTER

1. *Enumeration of General Properties.*—

- a. Indestructibility.
- b. Extension in space.
- c. Impenetrability.
- d. Mobility.
- e. Inertia.
- f. Divisibility.
- g. Porosity.
- h. Mutability of volume.
- i. Elasticity.

2. *Definition of Indestructibility.*—Matter is capable of changes in combination and molecular arrangements, but no particle of it can be created anew or destroyed. Care in observation traces the parts apparently missing. (Especially gases.)

a. Example of indestructibility: Burning candle under bell-glass; water and carbon dioxide formed and shown.

3. *Definition of Extension in Space.*—Three dimensions. Measurement by comparison with standards of unity. Space occupied called volume. English and American standards of longitudinal measure, the foot; French, the metre.

a. Relations of metric to the United States measures and weights: The unit of the measure of length is the metre, obtained by dividing the distance from the earth's equator to the pole into 10,000,000 parts. Divisions of the units of measure or weight are indicated by the prefix of Latin numerals; multiples by Greek.

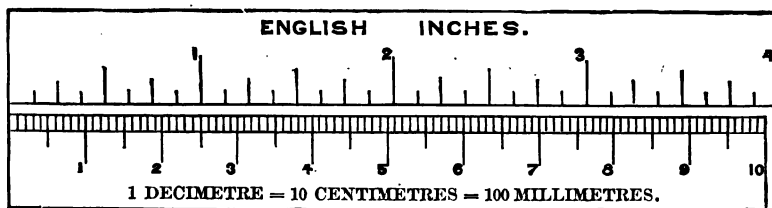


Fig. 1.

- 1 metre (M) = 10 decimetres (Dm.).
- = 100 centimetres (Cm.).
- = 1,000 millimetres (Mm.).
- 10 metres = 1 dekametre.
- 100 metres = 1 hektometre.
- 1,000 metres = 1 kilometre.

1 metre = 39·37079 inches.*
 1 decimetre = 3·93708 inches.
 1 centimetre = 0·39371 inches.
 1 millimetre = 0·03937 inches.
 1 inch = 25·3995 millimetres.
 1 cubic metre = 35·31658 cubic feet.
 1 cubic centimetre of water at its point of greatest density (4° C.) weighs 1 gramme, the unit of weight.
 1 gramme = 10 decigrammes = 100 centigrammes = 1,000 milligrammes; 10 grammes = 1 dekagramme; 100 grammes = 1 hektogramme; 1,000 grammes = 1 kilogramme.
 1 gramme = 15·43234874 grains troy.
 1 grain troy = 64·799 milligrammes.
 1 kilogramme = 2·6803 pounds troy = 2·20462 pounds avoirdupois.
 1,000 cubic centimetres (C. c.) = 1 cubic decimetre = 1 litre.
 1 cubic centimetre = 16·23 minims. 1 litre = 33·81 fluidounces = 2·1135 pints wine measure (or 1·76 pints imperial measure) = 61·027 cubic inches. 1 litre water weighs 1 kilogramme. 1 gramme per litre = 58·33 grains per wine gallon = 70 grains per imperial gallon.
 —[Curtman's Translation and Revision of Beilstein.

4. *Definition of Impenetrability.*—No matter can at the same time occupy the same space as other matter. Apparent exceptions explained.

a. Examples of impenetrability: (1) Air in test-tube inverted in water. (2) Solid immersed in liquid. (3) Diving-bell.

5. *Definition of Mobility.*—Change of position of a body by application of force.

a. Motion—a body in the act of changing place.

b. Velocity—the rate of changing place.

v (velocity) = $\frac{s \text{ (space)}}{t \text{ (time)}}$. It is expressed sometimes in terms of feet per second, miles per hour, or miles per second.

c. Table of velocities.

	Feet per second.	Miles per hour.
Man walking.....	4·40	3·00
" running.....	14·66	10·00
Horse trotting.....	40·00	27·00
Moderate wind.....	10·26	7·00
Storm.....	73·33	50·00
Sound.....	1,118·66	762·00
Rifle ball.....	1,466·66	1,000·00
		Miles per second.
Railway train.....		·02
Earth in motion around its axis point on equator.....		·29
Cannon ball.....		·44
Earth in its orbit.....		18·97
Light.....		185,500·00
Electricity.....		288,000·00

* The standard legalized by United States Congress, and by the British Parliament and Pharmacopœia, while the United States Pharmacopœia adopts Clarke's correction: 1 metre = 39·370432 inches.

6. *Definition of Inertia*.—Property of permanence in present state of rest or motion. Inability to stop or initiate motion unless from impulses from without.

a. Explanation: (1) A body at rest requires time for motion to be communicated to all parts. (2) A body in motion requires time to communicate the arresting force from one portion to another.

b. Examples of inertia: (1) Fracture of base of skull by *contre coup*. (2) Persons falling backward when standing in car suddenly started, and falling forward when it stops. (3) String attached to a door, when slowly pulled will close door, etc., but will break when quickly pulled. (4) Card-board thrown from under ball by sudden motion.

7. *Definition of Divisibility*.—Capacity of subdivision of matter into distinct parts.

a. Examples of divisibility: (1) Odor of musk. (2) Spray of ether. (3) Vapor of sodium coloring the flame yellow. (4) Expansion of gases. (5) Gold may be beaten out so thin that it will transmit green or blue rays of light. (6) One grain of gold can be subdivided into 95,000,000 particles that are visible under a microscope magnifying 1,000 diameters.

One five-hundred-thousandth of an inch is invisible under the best lenses made at the present time (1884).

(7) Markings of diatoms.

	Lines to the inch.
<i>Pinnularia nobilis</i>	12,500
<i>Navicula rhomboides</i>	75,000
<i>Pleurosigma angulatum</i>	50,000
<i>Amphipleura pellucida</i>	125,000

(8) Nobert's plates with over 120,000 lines to the inch. (9) Chemical precipitates: 1 part of the precipitate from a solution of nitrate of silver + a solution of chloride of sodium is visible in 138,000 parts of water, or 1 grain is visible in 200 fluidounces of water; a solution of acetate of lead + sulphuretted hydrogen; a solution of an iron salt + a solution of ferrocyanide of potassium, etc. (10) There are 32,000 red blood corpuscles to the linear inch.

b. Limit: The atom or the molecule.

8. *Definition of Porosity*.—The quality or state of having pores or interstices.

a. Examples of porosity: (1) Sugar. (2) The earth. (3) Lead. (4) Permeation of membranes by fluids (osmosis). (5) Gases pass through hot metals. (6) Diffusion of hydrogen, etc. (7) Solution of salts in water, etc. (8) Dialysis.

9. *Definition of Mutability of Volume*.—The properties of compressibility and expansibility.

a. Explanation: (1) Compressibility is that property of matter by

virtue of which it occupies less space when the pressure upon it is increased. (2) Expansibility is that property of matter by virtue of which it occupies more space when the pressure upon it is decreased.

b. Condensation of gases by compression: All gases were liquefied by Pictet and Cailletet in December, 1877.

c. Example of mutibility of volume: (1) Carthesian imps [Figs. 2 and 3]. (2) Pop-gun. (3) Pneumatic rifle.



FIG. 2.



FIG. 3.

d. Compressibility of liquids: Liquids possess very little compressibility. Water is compressible at 15° C. 0·0000457 for each atmosphere of pressure. Mercury is still less compressible.

e. Compressibility of solids: Metals may be compressed by hammering or by coining.

10. *Definition of Elasticity.*—The property of matter by virtue of which it resumes its original volume after compression or expansion.

a. Gases have the most elasticity.

b. Liquids have the least elasticity.

c. Solids vary greatly as to elasticity.

d. Example of elasticity: (1) Rubber ball. (2) Ivory ball. (3) Ricochet shooting in artillery practice.

The atoms are held together by chemical attraction (affinity), in groups called molecules. The molecules are held together by cohesion. Luminiferous ether fills the interstices between the molecules. The interstices may become greater by expansion, or less by compression. As a result of this ultimate constitution of matter, we have mutability of volume, porosity, and elasticity.

IV. CLASSIFICATION OF FORCES.

1. *Molar*.—

a. Mechanical motion.

b. Gravitation.

Cohesion and adhesion are molecular, and only indirectly molar.

2. *Molecular*.—

a. Cohesion.

b. Adhesion.

c. Heat, } Repulsive.

d. Light, }

e. Electricity, } Polar.

f. Magnetism, }

3. *Atomic*.—

a. Chemical affinity.

Another subdivision is: (1) Attractive forces. (2) Repulsive forces. (3) Polar forces.

V. GRAVITATION.

1. *Definition of Gravitation*.—The attraction of masses for each other.

a. Newton's Law: The amount of attraction of two bodies is proportional to the mass of each, and inversely proportional to the square of the distance.

2. *Definition of Weight*.—Weight is the effect produced by the attraction of the earth.

a. Standards and units of weight: (1) Avoirdupois, the grain. (2) Troy, the grain. (3) Metric, the gramme.

b. Absolute weight is the measure of the mass or quantity of matter.

c. Specific weight is the weight referred to a unit of comparison.

One cubic foot of water = 998.7 avoirdupois ounces. This is a standard not much used at present.

3. *Definition of Specific Gravity*.—The weight of a body referred to an equal bulk of a standard.

a. Water at its greatest density (4° C. or 39.2° F.) is the standard for solids and liquids.

b. Hydrogen is the standard for gases.

Air was formerly used. Air = 1.00, water = 773; water = 1.00, air = 0.0013.

4. *Manner of Ascertaining Specific Gravity of Solids*.—Ascertain the absolute weight, and then the weight of an equal bulk of water, and divide the absolute weight by the weight of the water.

a. Archimedes' Law: A body immersed in water loses weight equal to its volume of water (or any other liquid in which it may be immersed).

b. Example of ascertaining specific gravity of solids: Mohr's balance for taking specific gravity [Fig. 4].

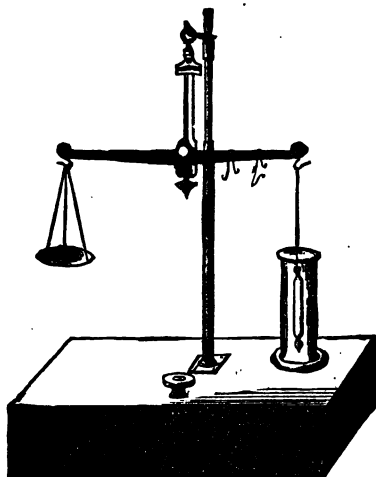


FIG. 4.

5. Manner of Ascertaining Specific Gravity of Liquids.—Ascertain either by weighing the liquid in a bottle whose tare and contents of distilled water have been ascertained, or by one of the various kinds of areometers.

a. Example of ascertaining specific gravity of liquids: Hydrometers and specific gravity bottles [Figs. 5, 6, 7, 8, 9, 10].



FIG. 5.



FIG. 6.



FIG. 7.



FIG. 8.



FIG. 9.

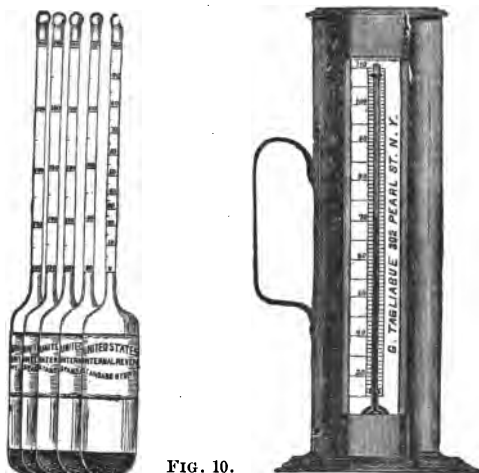


FIG. 10.

6. *Specific Gravity of Gases*.—There are several methods of determining the specific gravity of gases. Among them are Regnault's, Dumas', and Bunsen's. At present, hydrogen is the standard; formerly, air was used as the standard.

a. Gases compared with water at 0°C . and 760 millimetres mercurial pressure, according to Regnault:

Air.....	0.001293187
Nitrogen.....	0.001256157
Oxygen.....	0.001429802
Hydrogen.....	0.00089578

b. Weight of Gases: 1 cubic metre of air = 1,293.2 grams; 1 cubic metre of hydrogen = 89.6 grams; 1 cubic metre of carbon dioxide = 1,977.0 grams.

One gramme dry air at ocean level in 45° north latitude, at 0°C . and 760 millimetres mercurial pressure = 773.526 cubic centimetres.

One thousand cubic centimetres of oxygen at ocean level in $52^{\circ} 36'$ north latitude, at 0°C . and 760 millimetres mercurial pressure = 1.43028 grammes.

7. *Table of Specific Gravities*.—

a. Solids. Distilled water at 4°C . = 1.000.

Iridium.....	22.700	Palladium.....	11.800
Hammered platinum.....	22.100	Hammered lead.....	11.450
Melted platinum.....	20.860	Melted lead.....	11.350
Osmium.....	21.040	Ruthenium.....	11.400
Hammered gold.....	19.320	Silver.....	10.470
Melted gold.....	19.250	Bismuth.....	9.820
Uranium.....	18.400	Copper.....	8.880
Wolframium.....	17.600	Cast copper.....	7.790
Frozen mercury.....	14.000	Cadmium.....	8.650
Rhodium.....	12.100	Molybdenum.....	8.610

Brass	8.390	Fluor spar	3.150
Nickel	8.280	Marble	2.840
Manganese	8.000	Crystals of quartz	2.680
Steel	7.820	Aluminium	2.670
Cobalt	7.810	Strontium	2.540
Cast iron	7.21 to 7.790	Porcelain	2.490
Tin	7.290	Red phosphorus	2.080
Zinc	6.80 to 7.200	Rhombic sulphur	2.045
Antimony	6.710	Prismatic sulphur	1.982
Tellurium	6.110	White phosphorus	1.770
Crystalline arsenic	5.730	Magnesium	1.740
Iodine	4.950	Calcium	1.550
Amorphous arsenic	4.710	Rubidium	1.520
Barium	4.500	Sodium	0.970
Sulphate of barium	4.430	Ice	0.920
Diamond	3.520	Potassium	0.860
Flint glass	3.200	Lithium	0.590

b. Liquids. Distilled water at 4° C. = 1.000.

Mercury	13.598	Glycerin	1.260
Bromine	2.966	Milk	1.030
Sulphuric acid (H ₂ SO ₄)	1.848	Sea water	1.026
Chloroform	1.488	Distilled water	1.000
Nitric acid (HNO ₃)	1.500	Absolute alcohol	0.793
Bisulphide of carbon	1.272	Ether	0.715

c. Gases and vapors. Reduced to 0° C.

	Hydrogen =1.000.	Air =1.000.
Hydrogen	1.0000	0.06926
Carbon monoxide	13.9650	0.9670
Nitrogen	14.0200	0.97130
Oxygen	15.9633	1.10563
Chlorine	35.3700	2.45000
Carbon dioxide	21.9450	1.52020
Ammonia	8.5100	0.58900

d. The specific gravity of gases, with hydrogen as a standard, is very easily calculated by dividing the molecular weight by two. (To get the specific gravity compared to air this must be divided by 14.435, the specific gravity of air.)

8. *Definition of Center of Gravity.*—Direction of force of gravity on the earth towards its center. The line indicated by the plumb-line.

a. Examples of center of gravity: (1) Bodies of various forms. (2) Bodies with uniform density. (3) Bodies with un-uniform density, etc.

9. *Equilibrium.*—Every particle of a body being attracted, equilibrium (in suspended solids) will be reached when the center of gravity is supported, *i. e.*, support at the point of intersection of direct on of gravity in various positions.

a. Stable equilibrium exists when a body tends to return to its original position after it has been somewhat displaced.

b. Unstable equilibrium exists when a body tends to depart farther from its original position after it has been somewhat displaced.

c. Neutral equilibrium exists when a body remains at rest in any adjacent position after it has been displaced.

d. Examples of equilibrium: (1) Level surfaces of liquids. (2) Communicating tubes. (3) Fountains, etc. (4) Stratification of gases.

VI. MECHANICAL MOTION.

1. *Kinds of Motion*.—Absolute and relative; uniform and varied.

a. Uniform motion is that in which equal spaces are described in equal times:

b. Varied motion is that in which unequal spaces are described in equal successive units of time. It is when a body is *accelerated* and *retarded*.

2. *Rate of Motion*.—[See page 5.]

3. *Velocity of Motion*.—[See page 5.]

4. *Newton's Laws of Motion*.—

a. Every body continues in a state of rest, or of uniform motion in a straight line, unless acted upon by some external force.

b. Motion, or change of motion, is proportional to the force impressed, and is in the direction of the line in which that force acts.

c. Action and reaction are always equal, and are in opposite directions.

5. *Momentum*.—Is the mass of a moving body multiplied by the velocity (quantity of motion).

a. Large masses moving slowly and small masses moving swiftly may have the same momentum.

VII. VARIOUS MEMORANDA.

1. *Earth*.—

a. Measurements of the earth:

	English feet.	English miles.
Mean equatorial diameter.....	41,848,380	7,925.86
Polar diameter.....	41,708,710	7,899.33
Difference.....	139,670	26.47

The difference is about six German geographical miles.

An equatorial section of the earth would form an ellipse with about two miles difference in diameter—*i. e.*, a line through the earth at the equator, from 14° longitude east of Greenwich (Loango, in eastern Africa) to 165° 8' longitude west of Greenwich (Christmas Islands), is two miles longer than a line at right angles to it.

	Metres.
Equatorial mean diameter.....	12,754,796
Polar mean diameter.....	12,712,160
Meridian quadrant from equator to pole.....	10,000,000

This quadrant is as first measured by Delambre and Mechain. It gave a flattening of 1:292.

Former measurements by Bouguer and Condamine gave an axial difference of 1:292.

Later corrections of calculations show still less flattening, making the meridian quadrant 10,000,856 metres, and the axial flattening 1:299.

b. Shape of the earth: An oblate spheroid.

c. Velocity of rotation of the earth around its axis (Helmholz):

	Metres per second.
At equator (0°).....	463
At 45°.....	327
At 60°.....	231

d. Distance of the earth from the sun: About 91,000,000 miles.

Cook's measurements from the transit of Venus (1874), was 93,000,000, and from the transit of Mars 91,718,000 miles.

This distance is equal to 24,000 earth-diameters.

Earth's distance from the moon is equal to sixty earth-diameters.

To travel the distance of the earth from the sun, a cannon ball, going at the rate of 400 yards per second, would require 13·25 years, and the sound of the report would reach it six months later. A railway train at forty miles per hour would require 270 years to travel the distance.

2. *Sun*.—

a. Diameter of the sun is 853,380 miles (Herschel's measurement 882,000 miles); this is equal to 112 earth-diameters.

b. Volume of the sun is more than 1,200,000 earth-volumes.

c. Weight of the sun is 300,000 earth-weights.

VIII. PNEUMATICS.

1. *Definition of Pneumatics*.—Pneumatics treat of the mechanical properties of aeriform bodies.

a. Aeriform bodies or gases are bodies whose molecules are kept separated by some repulsive force, so that no cohesion can occur between them.

b. Gases and vapors are divisions of aeriform bodies: (1) Gases remain aeriform at all ordinary natural temperatures. (2) Vapors are only aeriform at a somewhat elevated temperature and are ordinarily known in liquid (or solid) form.

c. Permanent and coercible gases is an old distinction that can not be made now, for all gases have been liquefied. [See b, page 7.]

d. Atmospheric air is used to exemplify the mechanical properties of all gases.

2. *Air*.—

a. Composition of the air:

	Parts by weight.	Parts by volume.
Oxygen.....	23·10	20·80
Nitrogen	76·90	79·20

The air also contains small quantities of aqueous vapor, carbon dioxide, ammonia, etc.

b. Height of the air: The atmosphere forms an envelope around the earth, about forty-five miles thick, or about $\frac{1}{175}$ of the diameter of the earth. A globe eight inches in diameter, with an atmosphere one-twenty-fourth of an inch thick, would be in about the same proportion as that of the earth.

c. Density of the air: Gravitation causes a difference in the density of the strata, decreasing from bottom to top. One cubic inch of air taken from the bottom to the top of the atmosphere would expand to 12,000 cubic inches.

At earth's surface.		Miles high.
1 volume will expand to	2 volumes at.....	2.705
1 " " " "	4 " " " ".....	5.410
1 " " " "	8 " " " ".....	8.115
1 " " " "	16 " " " ".....	10.820
1 " " " "	64 " " " ".....	16.230
1 " " " "	128 " " " ".....	18.935
1 " " " "	256 " " " ".....	21.640

—[GRAHAM.]

d. Temperature of the air: The heat of the atmosphere diminishes in ascent. The following figures are only approximate:

An ascent of 350 feet =	1° F.
An ascent of 630 feet =	1° C.
An ascent of 192.02 metres =	1° C.

In Glaisher's ascent in a balloon from Greenwich, September 5, 1862, the following temperatures were observed:

Temperature.	Metres.	Miles.
15° C.....	Earth's surface.	
5° C.....	1600	1
0° C.....	3218	2
-13° C.....	6437	4
-19° C.....	8000	5
-24° C.....	11000	7

e. Line of perpetual congelation: This is the line of eternal snow. It varies about as follows:

	Feet high.
Under equator.....	15,000
At 60° latitude.....	3,200
At 76° latitude.....	1,000

f. Cause of diminution of heat: (1) The less dense atmosphere above absorbs less heat. (2) The greatest part of the heat is communicated to the air by contact with the earth's surface, this diminishes with the expansion of the atmosphere.

g. Weight of the air: (1) Air is matter affected by gravity, has weight and offers resistance. (2)

Cubic metre.	Grammes.		Units or weight.
1.....	89.6	hydrogen =	1.000
1.....	1,293.0	air =	14.442
1.....	1,000,000.0	water =	11,167.000
1.....	13,598,000.0	mercury =	151,803.000

(3) One cubic foot of air weighs 1.2 avoirdupois ounces. (4) A room 14x20x50 feet contains about 1,000 pounds of air. (5) One avoirdupois pound of air = 13.333 cubic feet.

h. Discovery of the weight of the air: Evangelista Toricelli (born 1608, died 1647), a pupil of Galileo, and professor of mathematics and philosophy at Florence, in 1643 made the celebrated experiment in proof of the weight of the atmosphere. (Occasioned by the failure of the suction pumps in a deep well.) He shortened the tube by using mercury instead of water.

Mercury is 13.598 times heavier than water; 11,000 times heavier than air; 151,803 times heavier than hydrogen.

Thirty inches of mercury is the counterpoise of the atmosphere, which, if of uniform density as at bottom, would reach 27,400 feet ($5\frac{25}{66}$ miles) in height.

i. The height of the mercury in barometers varies with the fluctuations of the atmosphere; it averages 760 millimetres, or 30 inches.

j. Above the mercury in the tube is the Toricellian vacuum.

k. Verification of Toricelli's discovery was made by Pascal, of Clermont, at whose instance Perrier ascended *Puy le Dome* 5,000 feet above the ocean level (2,830 feet above bottom station), and found the mercury column three inches lower than at the bottom.

Verifications have also been made by balloon ascension by Glaisher and others.

l. Measurement of the height of the atmosphere by barometers:

At 50° F.	Height.
30 inches =	sea level.
29 inches =	917 feet.
28 inches =	1,860 feet.
27 inches =	2,830 feet.
26 inches =	3,830 feet.
25 inches =	4,861 feet.
20 inches =	11,000 feet.
10 inches =	27,500 feet.

In round numbers, every inch is equal to 900 feet. Corrections must be made for temperature and fluctuations.

m. Varieties of barometers: (1) Syphon [Fig. 11]. (2) Fountain. (3) Ordinary. (4) Aneroid [Fig. 12].

n. Calculation of the weight of the atmosphere: A mercury column 30 inches high, with a base of 1 square inch is equal to 30 cubic inches of mercury, and weighs 14.6 avoirdupois pounds, hence the pressure of the atmosphere is equal to 14.6 avoirdupois pounds per square inch

One cubic inch of mercury weighs 3,408.183 grains. Water being lighter, as 1:13.598, will be counterbalanced about 33.83 feet high.

Each square foot of surface sustains 2,103.336 avoirdupois pounds

of atmospheric pressure when the mercurial barometer indicates thirty inches.



FIG. 11.



FIG. 12.

The pressure of the atmosphere on the human body is about twenty tons.

Respiration is affected by difference in air pressure on mountains, in deep wells, etc.

The entire weight of the atmosphere is about equal to a globe of copper sixty-three miles in diameter.

o. Effects of atmospheric pressure: Pumps explained. The action of suction pumps [Fig. 13] is due to atmospheric pressure, and not to suction as the name implies. They can not raise water above thirty-three feet. Force pumps [Fig. 14] have alternate action of suction and pressure.

p. Experiments and illustrations of atmospheric pressure, without air pump: (1) Toricelli's tubes. (2) Water in inverted cylinders. (3) Pipette [Fig. 15]. (4) Syphon [Fig. 16]. (5) Magic can and similar devices. (6) Tantalus cup [Fig. 17]. (7) Reservoir lamps, fountain inkstands, etc. (8) Cupping glasses.

q. The air pump [Fig. 18] was invented by Otto V. Guericke, in 1650, at Magdeburg; it was exhibited at Regensburg in 1654. Boyle constructed a similar pump shortly afterwards.

Guericke's hemispheres; which were exhibited before Emperor Ferdinand III., were twenty-two inches in diameter and twenty-four horses did not pull them apart.

r. Varieties of air pumps explained: (1) Single-acting. (2) Double-

acting. (3) Rotary. (4) Geisler's. (5) Sprengel's. (6) Bunser's water-air pump, etc.



FIG. 13.

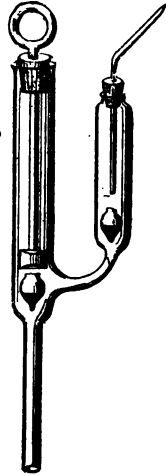


FIG. 14.



FIG. 15.

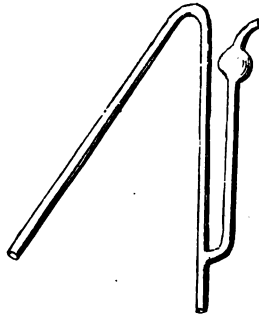


FIG. 16.



FIG. 17.

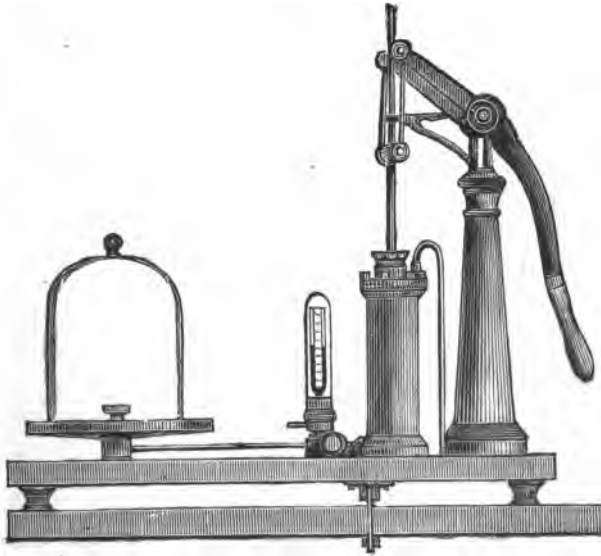


FIG. 18.

s. Experiments and illustrations of atmospheric pressure, with air pump (pressure against vacuum): (1) Philosopher's handcuffs. (2) Rubber on glass. (3) Bladder on glass. (4) Magdeburg hemispheres [Fig. 19]. (5) Fountain in vacuo [Fig. 20]. (6) Torricelli's tube in vacuo [Fig. 21]. (7) Fall in vacuo, as in guinea and feather tube. (8) Expansion of gases in vacuo, as beer foam, air closed in balloon, etc. (9) Candle extinguished in vacuo. (10) Boiling water under reduced pressure. (11) Cloud of aqueous vapor on rarefaction of air, caused by refrigeration. (12) Water hammer [Fig. 22], with or without air pump. (13) Balloon with hydrogen rises in air. (Charles, 1783, Montgolfier, Pilatre de Rossier, etc.)

t. Compression of air: (1) Condensing pumps explained [Fig. 23]. (2) Air gun. (3) Fountain jet, etc. (4) Fire engine air-chamber. (5) Hero's fountain (210 B. C.) [Fig. 24]. (6) Carthesian imps [Figs. 2 and 3, page 7]. (7) Pressure gauge [Fig. 26].

u. Elasticity of gases—Mariott's or Boyle's law (1660): The volume of a confined mass of gas is *inversely* proportional to the pressure imposed upon it, *i. e.*, the larger the pressure the smaller the volume, and the smaller the pressure the larger the volume.

IX. KINETIC THEORY OF GASES.

The molecules of a gas are in constant motion; the velocity increases with the temperature.

Their impulses against the walls of confining vessels constitute their pressure or tension.

Hydrogen molecules at 0° C. have a velocity of 6,097 feet per second.

The velocities of different gases are inversely proportional to the square root of their densities.

The more a gas is compressed the more molecules are moving in the same volume, and hence the greater the pressure.

X. COHESION.

1. *Definition of Cohesion.*—The attraction between the molecules of homogeneous matter.

The homogeneous matter may be either simple or compound.

a. Cohesion manifests itself most strongly in solids, less in liquids, and is entirely absent in gases.

b. Cohesion is only exerted at exceedingly small distances.

c. Cohesion shows itself by resistance to separating forces, and this resistance may be different, according to the mode of application of the separating force.

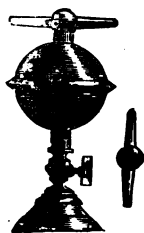


FIG. 19.

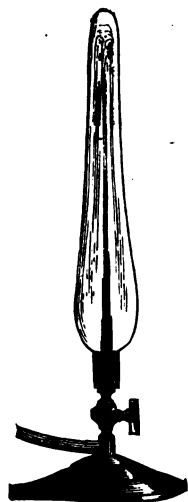


FIG. 20.



FIG. 21.



FIG. 22.



FIG. 23.

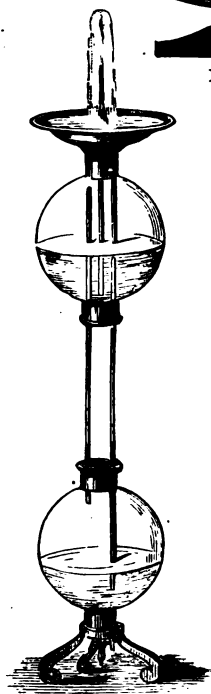


FIG. 24.



FIG. 25.



FIG. 26.

d. Longitudinal application of separating forces: (1) Compressing. (2) Pulling. (3) Crushing. (4) Tearing.

e. Transverse application of separating forces: (1) Bending. (2) Breaking. (3) Torsion. (4) Wrenching.

f. Mixed application of separating forces: Distortion, as the insertion of a chisel or wedge, the use of shears, etc.

g. We distinguish various forms of cohesion: (1) Hardness. (2) Brittleness. (3) Malleability. (4) Ductility. (5) Tenacity.

h. Definition of hardness: It is the resistance of a substance to abrasions, as filing or scratching.

i. Scale by which the hardness of minerals are measured:

Talc.....	1.0	Scapolite (Hornblende).....	5.5
Gypsum	2.0	Feldspar.....	6.0
Mica	2.5	Quartz.....	7.0
Calcsp.	3.0	Topaz.....	8.0
Fluorspar	4.0	Corundum.....	9.0
Apatite.....	5.0	Diamond.....	10.0

j. Definition of brittleness: The ease with which a body is fractured by compressing, pulling or bending it.

Brittleness is opposed to malleability, ductility, and tenacity.

k. Definition of malleability: The property of a body by which it yields under strokes of the hammer or other compression.

l. Definition of ductility: The property of a body by which it yields to a pull, combined with lateral compression, as the drawing of a metal into wire.

m. Definition of tenacity: The property of a body by which it resists a pull, wrench, or compression.

n. Heat is antagonistic to cohesion.

o. Cohesion of liquids.

p. Formation of drops.

q. Examples of cohesion of liquids: (1) A large globular drop of oil is formed in alcohol of the same specific gravity. (2) Tomlinson's cohesion figures with oils upon water. (3) Plateau's experiments with soap films, on frames. (4) Contraction of round soap bubbles (5) Ascension of soap films in funnel tube.

XI. ADHESION.

1. *Definition of Adhesion.*—The attraction between *unlike* bodies (at small distances).

2. *Examples of the Adhesion of Solids to Solids.*—

a. Friction.

b. Chalk to black-board.

c. Ice to other solids.

d. Amalgam to glass.

e. Hardened liquids.

3. *Examples of the Adhesion of Liquids to Solids.*—

- a. Glue.
- b. Cement.
- c. Water to solids.

4. *Example of the Adhesion of Gases to Solids.*—A dirty glass when filled with carbonic acid water, shows gas bubbles adhering to the rough or dirty spots.

A smooth, perfectly clean glass, remains free from bubbles.

5. *Effects of the Balancing of Cohesion and Adhesion.*—Different forms of surface of mercury or water in tubes, or by the immersion of glass tubes in mercury or water.

6. *Definition of Capillary Attraction.*—The manifestation of adhesion between solids and liquids in small capillary spaces.

a. The smaller the space occupied by the liquid, the greater the influence of the adhesion to the approximated solid.

b. Examples of capillary attraction: (1) Bundle of capillary tubes dipped into colored liquids. (2) Capillary tubes of different diameters. (3) Capillary spaces between two glass plates; if separated on one side a curve is produced in the surface of the liquid. (4) Blotting paper. (5) Lampwicks. (6) Sugar. (7) Efflorescence of crystallizing solutions by capillary attraction.

XII. CRYSTALLOLOGY.

1. *Crystallogology is divided into*—

- a. Crystallography.
- b. Crystallophysics.
- c. Crystallochemistry.

2. *All inorganic bodies are either amorphous or crystalline.*

3. *Definition of Amorphous Bodies.*—Those bodies which have no essential regular forms. They have equal cohesion in all directions, and break with a conchoidal fracture. They never have double refraction.

4. *Examples of Amorphous Bodies.*—

- a. Glass.
- b. Rosin.
- c. Rubber, etc.

5. *Definition of Crystalline Bodies.*—Those bodies which are bordered by plane surfaces, and have some regularity or symmetry in their forms. They have unequal cohesion in different directions which permits cleavage. They often have double refraction of light.

6. *When the same body exists in both the amorphous and the crystalline state, the two differ by the amorphous showing—*

- a. Less specific gravity.
- b. Less hardness.
- c. Lower melting point.

d. Different color.

e. More easily influenced by chemical reagents, solvents, etc.

7. *Crystal Systems*.—The lines of direction according to which the molecules are built up are called axes. They meet in a common center and are distinguished according to their number and position.

a. The regular system has three axes cutting each other at right angles. They are of equal dimensions.

b. The quadratic system has three axes cutting each other at right angles. Two of them are of equal and one of unequal dimension.

c. The hexagonal system has four axes. Three equal axes cut each other so as to make six angles of 60° , the fourth axis standing perpendicular to the plane of the three others.

d. The rhombic system has three axes cutting each other at right angles. They are all of unequal dimensions.

e. The monoklinometric system has three axes. Two of them are at right angles to each other, and the third is inclined to one of the two which are at right angles to each other. They are all of unequal dimensions.

f. The diklinometric system has three axes. Two of them are at right angles to each other and the third is inclined to both of them. They are all of unequal dimensions.

g. The triklinometric system has three axes. They are all inclined to each other and none of them at right angles. They are all of unequal dimensions.

8. *Forms are holohedric or hemihedric.*

9. *Twin crystals.*

10. *Distorted forms.*

11. *Regular System*.—This system has all of the forms developed equally in three directions at right angles to each other.

In drawing it is conventional to make aa [Fig. 27] equal to $a'a'$, and a'' a'' $\frac{1}{3}$ as long.

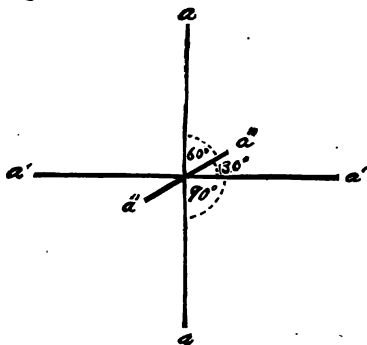


FIG. 27.

Letting " c " represent the point at the intersection of aa and $a'a'$, the angles aca' , aca'' , and $a'ca''$ are all right angled, but to represent perspective $a'ca''$ is drawn at an angle of 30° .

12. *Holoedric Forms of the Regular System.*—There are 7 holoedric forms in the regular system.

a. Octohedron [Fig. 28]— $a : a : a$, or O .

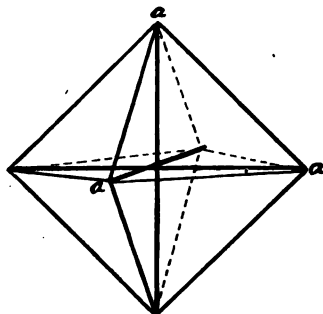


FIG. 28.

The bordering planes cut the axes at equal distances. It has 8 facets, which are equilateral triangles. Also 6 corners, in which 4 triangles meet. It has 12 edges.

This form occurs in: (1) Alum series. (2) Arsenic. (3) Arsenic trioxide. (4) Borax ($\text{Na}_2\text{B}_4\text{O}_7 + 7\text{H}_2\text{O}$). (5) Cadmium. (6) Cuprous oxide (Cu_2O). (7) Diamond. (8) Galena (PbS). (9) Gold. (10) Iron. (11) Lead. (12) Magnetic iron ore. (13) Magnesium. (14) Mercury. (15) Phosphorus. (16) Sodium. (17) Spinell, etc.

b. Cube or Hexahedron [Fig. 29]— $a : \infty a : \infty a$, or ∞O .

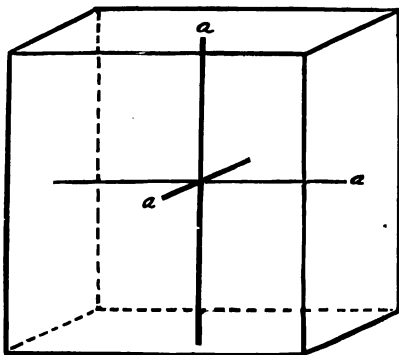


FIG. 29.

The bordering planes are squares, cutting one axis at right angles and parallel with each of the others. It has 6 square facets and 8 corners, in which 3 facets meet. Also 12 edges, in which 2 facets meet at right angles.

The following occur in this form: (1) Ammonium chloride. (2) Arsenical pyrites. (3) Calcium fluoride, CaF_2 (Cleavage O). (4) Copper. (5) Galena (Pb S). (6) Gold. (7) Iron pyrites. (8) Potassium. (9) Potassium bromide. (10) Potassium chloride. (11) Potassium iodide. (12) Palladium. (13) Silver. (14) Sodium chloride, etc.
c. Rhombic Dodekahedron [Fig. 30]— $a : a : \infty a = \infty O$.

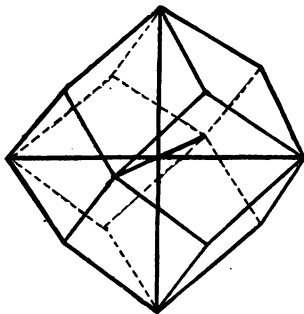


FIG. 30.

This form has 12 rhombic facets, cutting 2 axes at equal distances, and parallel with the third axis. It has 14 corners; 6 of them are at the meeting points of 4 facets, with their acute angles, and 8 corners at the point of meeting of 3 facets with their obtuse angles.

The following occur in this form: (1) Ammonium chloride. (2) Copper. (3) Diamond. (4) Garnet. (5) Phosphorus. (6) Silver. (7) Zinc blende (Zn S), etc.

d. Icositetrahedron [Fig. 31]— $a : m a : m a$, or $m O m$ (m being a larger number than 1, either 2 or 3 as $2 O 2$), $2 O 2 = \text{Leucitohedron}$; $3 O 3 = \text{Leucitoïd}$.

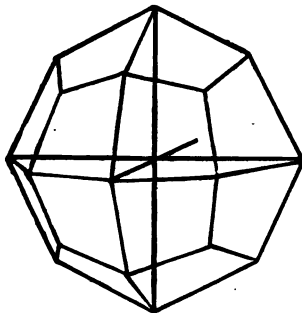


FIG. 31.

This form has 24 trapeze facets which are also called deltoïds. It has 26 corners and 48 edges. Of the latter 24 are shorter and 24 longer.

The following occur in this form: (1) Analcim. (2 O 2). (2) Garnet. (3) Gold (3 O 3). (4) Leucit. (5) Silver, etc.

e. Triakisoctohedron [Fig. 32]— $a : a : m a = m O (2 O)$.

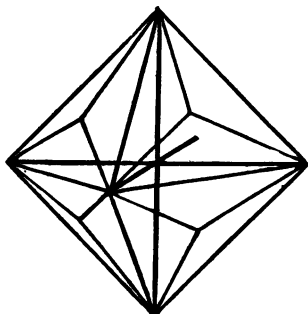


FIG. 32.

This form has 24 isoscelos triangular facets, resembling three-sided pyramids placed on octohedron.

It has 14 corners; 6 of them axial at meeting of 6 facets, 8 are hexahedral at meeting of 3 facets. It has 36 edges.

The following occur in this form: (1) Boracite. (2) Chromic iron. (3) Copper. (4) Diamond. (5) Garnet. (6) Gold. (7) Phosphorus. (8) Silver. (9) Silver amalgam. (10) Spinell, etc.

f. Tetrakisohexahedron [Fig. 33]— $a : m a : \infty = m O \infty$, $m = 2$, or $= 3$, or $= 3/2$, or $= 5$, etc. (2 O ∞).

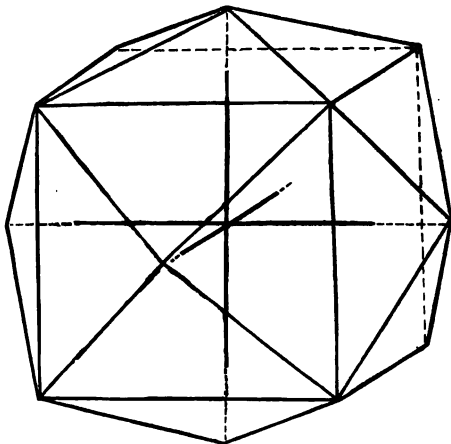


FIG. 33.

This form has 24 isoscelos triangular facets.

It has 14 corners; 6 of them are axial (octohedron) and 8 hexadral. It has 36 edges.

The following occur in this form: (1) Calcium fluoride. (2) Copper. (3) Diamond. (4) Iron pyrites. (5) Sodium chloride, etc.

g. Hexakisoctohedron [Fig. 34]— $a : m a : n a = m O n, 3 O \frac{2}{3}, 3 O \frac{1}{3}, 2 O 4$, etc.

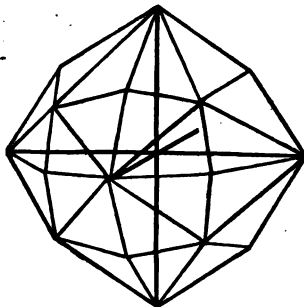


FIG 34.

This form has 48 triangular facets.

It has 26 corners, 6 of them axial (octohedral) with 8 facets to each corner, 8 hexahedral with 8 facets to each corner, and 12 symmetric with 4 facets to each corner. It has 72 edges.

The following occur in this form: (1) Diamond. (2) Iron pyrites, etc.

13. Hemihedric Forms of the Regular System.—

a. Tetrahedron, $+\frac{O}{2}$, or $-\frac{O}{2}$.

This form has alternate facets of octohedron enlarged to disappearance of the intermediate ones.

Often both $+$ and $-$ are present (Schlippe's salt).

The cube forms no hemihedric form, as it has only 6 facets, while for a hemihedric form the number of facets must be divisible by 4 at least. Neither is there any hemihedric form known of the rhombic dodecahedron.

b. The triakisoctohedron forms a hemihedric crystal in which every alternate octant of 3 facets grows.

It is called deltoïd-dodecahedron. Fahlerz and zinc blende are examples, but it does not occur on artificial crystals.

c. The triakistetrahedron is the hemihedric form of icositetrahedron. It has the alternate octants enlarged.

d. The pentagonal dodecahedron is the hemihedric form of the tetrakisbexahedron, by the enlargement of alternate facets.

e. The hexakisoctohedron forms 2 hemihedra: (1) By enlargement of alternate octants, as in the hexakistetrahedron in *O m*. (2) By enlargement of alternate pairs of facets, as in the trapezoid icositetrahedron, $\frac{m O n}{2}$.

14. *These forms may all occur singly or combined, so that several polyhedral and hemihedral forms may concur in the same crystal.*

15. *Quadratic System.*—*

a. The following occur in this system: (1) Tin, metallic and the native oxide (Sn O_2). (2) Rutil, TiO_2 . (3) Copper pyrites. (4) Red iodide of mercury. (5) Phosphate of potassium. (6) Sulphate of nickel. (7) Hyacinth. (8) Zircon. (9) Apophyllite. (10) Cyanide of mercury. (11) Ferrocyanide of potassium, and many others.

16. *Hexagonal System.*—*

a. The following occur in this system: (1) Graphite. (2) Mica. (3) Arsenic. (4) Tellurium. (5) Antimony. (6) Bismuth. (7) Quartz. (8) Corundum. (9) Sapphire. (10) Calcite. (11) Cinnabar. (12) Dolomite spar. (13) Chabasite. (14) Tourmaline, and many others.

17. *Rhombic System.*—*

a. The following occur in this system: (1) Iodine. (2) Sulphur. (3) Nitrate of potassium. (4) Pyrolusite. (5) Yellow iodide of mercury. (6) Corrosive sublimate. (7) Chloride of barium. (8) Chromate of potassium. (9) Manganate of potassium. (10) Sulphate of potassium. (11) Arragonite. (12) Permanganate of potassium. (13) Sulphate of barium. (14) Sulphate of strontium. (15) Carbonate of sodium ($+ 7\text{H}_2\text{O}$). (16) Citric acid. (17) Nitrate of silver. (18) Rochelle salts, and many others.

18. *Monoclinometric System.*—*

a. The following occur in this system: (1) Sulphur (prismatic). (2) Selenium. (3) Carbonate of sodium ($+ 10\text{H}_2\text{O}$). (4) Borax ($\text{Na}_2\text{B}_4\text{O}_7 + 10\text{H}_2\text{O}$). (5) Chlorate of potassium. (6) Phosphate of sodium. (7) Sulphate of calcium. (8) Iron. (9) Manganese. (10) Cadmium. (11) Cobalt. (12) Magnesium. (13) Hornblende. (14) Augite. (15) Oxalic acid. (16) Tartaric acid. (17) Ferri-cyanide of potassium. (18) Sugar. (19) Piperine, and many others.

19. *Diclinometric and Triclinometric Systems.*—*

a. The following occurs in these systems: (1) Albite. (2) Sulphate of copper. (3) Boric acid. (4) Dichromate of potassium, and many others.

XIII. UNDULATIONS.

1. *Explanation of Undulations.*—A body may be set in motion as a whole without relative change of situation of its molecules, or else it

*For description of these systems see page 22.

may be moved in vibrations, waves, or undulations, which change the relative position of its molecules.

a. Difference in the vibrations in solids, liquids, gases, or ether.

2. *Form of Wave in Elastic Cord or in Liquid (exaggerated)* [Figs. 35 and 36].—

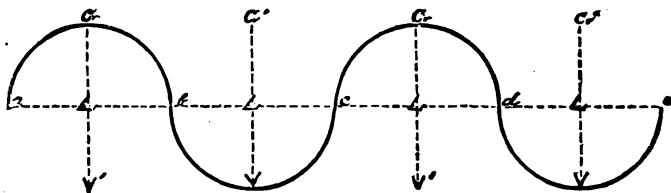


FIG. 35.

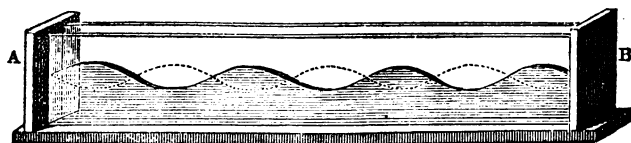


FIG. 36.

- a. The line of level in quiescence is LL.
- b. The wave crests or elevations are Cr Cr.
- c. The wave valleys or depressions are V V.
- d. The wave height is L Cr.
- e. The wave depth is L V'.
- f. The wave amplitude is the height L Cr plus the depth L V', or Cr V'.
- g. The phase of elevation is a b or c d.
- h. The phase of depression is b c or d e.
- i. The wave length is the phase of elevation a b, plus the phase of depression b c, or a c.
3. *Progressive undulations are not real forward movements.*
4. *Longitudinal, transverse, and torsional undulations.*
5. *Nodes.*—The points of intersection of undulations with level at a, b, c, etc, are called nodes.
6. *Examples of Nodel Lines in Vibrating Plates.*—Chladni's figures [Fig. 37].
7. *Longitudinal Vibration.*—The longitudinal vibrations in air and other gases are alternate condensations and rarefactions.
8. *Reflected waves.*
9. *Interference and combination.*
10. *Sound Waves.*—The sound waves in air or other gases are composed of condensation and rarefaction phases.



FIG. 37.

a. The rapidity and size of waves governs the *pitch and tone* of the sound.

b. The amplitude or degree of condensation governs the *intensity* of sound.

c. The intensity of sound is in *inverse* proportion to the *square* of the distance from the sound-producing body.

d. The velocity of sound varies according to the medium in which it travels.

At 0° C. it is in:

Air equal to.....	1,090 feet in 1 second
Oxygen equal to.....	1,040 " 1 "
Carbon Monoxide equal to.....	1,107 " 1 "
Hydrogen equal to.....	4,164 " 1 "
Carbon Dioxide equal to.....	858 " 1 "
Nitrogen Monoxide equal to.....	859 " 1 "

e. The velocity of sound in liquids and solids is greater than in air.

At 0° C. it is in:

Water equal to.....	4,700 feet in 1 second
Lead equal to.....	4,030 " 1 "
Silver equal to.....	5,717 " 1 "
Steel equal to.....	16,600 " 1 "
Glass equal to.....	16,600 " 1 "

f. Waves of equal intensity meeting in opposite phases, extinguish each other.

g. An *echo* is the result of the reflection of a sound wave.

11. *Musical Sound Waves.*—

a. The highest audible sound has 24,000 vibrations in one second of time. Its wave length is 0.545 inches.

b. The lowest audible sound has 16 vibrations in a second of time. Its wave length is 68.125 feet.

It is called C — 3. Its octave C — 2 makes 32 vibrations per second, and has a wave length of 34.0625 feet.

C — 1 makes 64 vibrations per second. Its wave length is 16.53125 feet.

C makes 128 vibrations per second. Its wave length is 8.765 feet.

C_1 makes 256 vibrations per second. Its wave length is 4.3825 feet..

C_2 makes 512 vibrations per second. Its wave length is 2.191 feet.

b. Table of vibrations:

	C	D	E	F	G	A	B	C
Relative number	1	$\frac{9}{8}$	$\frac{5}{4}$	$\frac{4}{3}$	$\frac{2}{3}$	$\frac{5}{6}$	$\frac{15}{8}$	2
Number per second.....	128	144	160	170 $\frac{2}{3}$	192	213 $\frac{1}{3}$	240	256
Wave length, in feet.....	8.7	7.7	7	6.5	5.8	5.2	4.6	4.4

XIV. LIGHT.

1. Nature of Light.—

a. Various hypotheses: Emanation—Newton (1666), Huyghens (1690). Undulation theory—Euler, Young (1802), Fresnel (1815), etc.

b. Light is now considered to be a vibratory mode of motion of the *luminiferous ether*, consisting of waves of a definite length (between $\frac{1}{37640}$ and $\frac{1}{38750}$ inch) and of definite velocity.

c. Those waves which have a larger size are heat rays and those of less size, actinic waves.

d. The heat waves extend up to 976° F., and light begins at 977° F.

2. *Definition of Light.*—Light is an undulatory motion of the luminiferous ether, comprising those waves which are of a size and velocity suitable to affect the eye.

3. *Source of Light.*—Bodies whose molecules are in a state of rapid vibration communicate their vibrations to the luminiferous ether, and are luminous when their vibrations are of the proper magnitude.

a. Such vibrations may be produced by: (1) Mechanical motion. (2) Heat. (3) Electricity. (4) Chemical affinity. (5) In fact by the transformation of any form of energy.

b. The sun and stars are self-luminous.

c. Opaque bodies reflect light.

4. Experiments with Light.—

a. Drummond's calcium light.

b. Luminous and non-luminous gas flame of Bunsen's burner [Fig. 38]. (1) The Bunsen burner consists of a small burner A, inclosed in a tube B, open on top E, to which air is admitted by round openings near the bottom of the tube. By turning a perforated collar C, whose holes correspond with those in the tube which it surrounds, the admission of air may be regulated or shut off entirely. The supply of coal gas is admitted by the lateral tube D, which is connected with the gas fittings of the laboratory by a rubber tube, and is regulated by an ordinary stopcock. (2) The following parts of the flame [Fig. 39] are distinguished: A *dark cone* in the center, rising from the base, consisting of unburnt gas mixed with a considerable portion of air. The outer flame consists of burning carbon monoxide, in which different parts are distinguished. The *zone of fusion* (Z F) near the upper end of the dark cone, in which the temperature attains

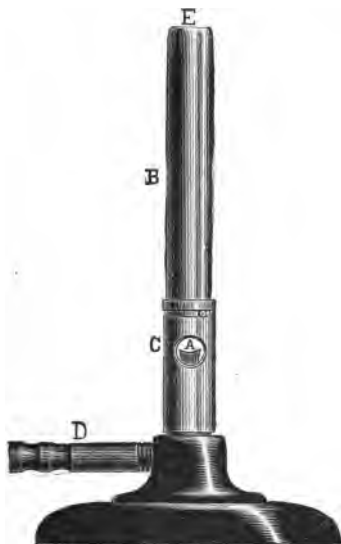


FIG. 38.*

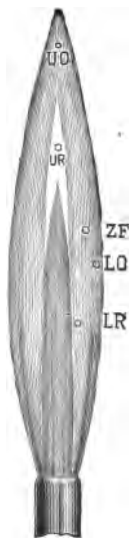


FIG. 39.*

its highest degree half way between the dark cone and the outer margin. The *upper* (U O) and *lower oxidation* (L O) flames, which, being near the end of the flame, are richest in oxygen, and are the places in which oxidation occurs most readily, and are used for the oxidation of metals, etc. The *upper* (U R) and *lower reduction* (L R) flames, being situated in the inner part of the flame, containing most carbon are employed for reducing metallic oxides, etc. The upper reduction flame (U R) is produced by partially shutting off the air holes, and forms a luminous tip of the dark cone. The luminous flame is surrounded by a non-luminous margin or seam, called the *mantle*, in which more perfect combustion occurs by contact with the oxygen of the surrounding air. This mantle, though not so readily observed in the non-luminous flame, nevertheless forms even there a separate part of the flame, and may be shown by coloration with boric acid or other substance, to be sharply differentiated from the inner part of the non-luminous flame, which lies between the mantle and the dark cone. (3) For blow-pipe operations and the bending of glass tubes, flat jets [Fig. 40] of different sizes may be attached to the top of the burner so as to give a thin, flat flame of such breadth as is most suitable to the work to be done. Care must be taken that the flame does not flash back and burn at the end of the little jet

*From Curtman's Chemical Analysis.



FIG. 40.*

[A, Fig. 38]. Heating of the lamp and spoiling of the rubber connections results from this; the flame at the top E, has then a peculiar odor and appearance, which at once reveals the accident. The lamp must in this case be extinguished and relighted. (4) The blow-pipe flame [Fig. 41].

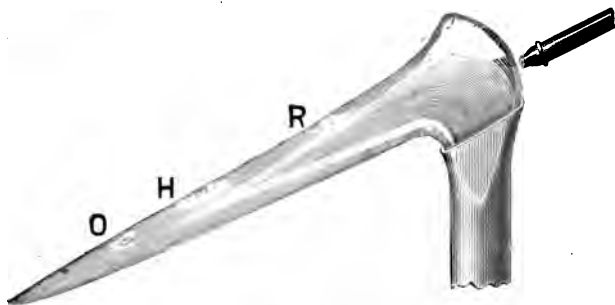


FIG. 41.*

- c. Platinum heated becomes luminous.
- d. Light produced by chlorochromic acid in hydrogen flame.
- 5. *Homogeneous Bodies*.—Light is transmitted in homogeneous media in a straight direction.
- 6. *Transparent Bodies*.—There is a more or less perfect passage of light through transparent bodies.
- 7. *Translucent Bodies*.—There is a very imperfect or only partial passage of light through translucent bodies.
- 8. *Opaque Bodies*.—There is no transmission of light at all through opaque bodies.
- 9. *Luminous Bodies*.—Luminous bodies are considered as an aggregate of luminous points.
 - a. Vibrations from luminous points proceed in all directions.
 - b. A single line is called a ray.

*From Curtman's Chemical Analysis.

c. A collection of contiguous rays forms a *pencil of rays*.

d. A pencil of light may be *divergent*, *convergent*, or *parallel*.

e. The parallel pencil is also called a *beam of light*.

10. *Velocity of Light*.—The velocity of light was first observed in 1675 by Olaf Roemer on the satellites of Jupiter.

a. The difference of time in the eclipses in the different positions of the earth is equal to 987.5 seconds (16 minutes, 27.5 seconds). As this difference is equal to the distance of double the distance of the earth from the sun ($2 \times 91,718,000$ miles), it will give 8 minutes, $13\frac{3}{4}$ seconds for the time required for light to reach the earth from the sun, and gives the velocity of light at 186,300 miles per second of time.

b. Captain Cook (1769), from the transit of Venus, calculated the distance of the sun from the earth as being 95,000,000 miles.

11. *Direct Measurement of Velocity of Light*.—According to Fizeau the velocity of light is 313,274,304 (194,664.9 English miles) metres per second of time. According to Foucault it is 298,000,000 metres (185,172 English miles) per second.

Sound travels 1,090 feet per second.

Electricity travels 200,000 miles per second.

12. *Shadows*.—

a. When light proceeds from a point [A, Fig. 42] with an opaque

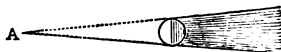


FIG. 42.

body intercepting, it will cast a shadow increasing in size with the distance from the opaque body. [Also see Fig. 43.]



FIG. 43.

b. When the source of light is larger than a point, especially when

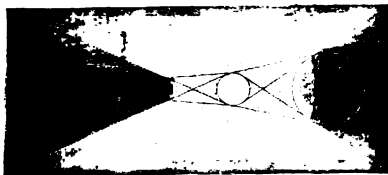


FIG. 44.

it is larger than the opaque body [Fig. 44], then each pencil of rays will cast an independent shadow. The co-operation will produce the *true shadow*, or *umbra* and the *penumbra*.

13. *Images Formed by Direct Light.*—

a. Image is inverted.

b. If the opening in a screen is large, the form will be indistinct with juxta and superposition of many images. (1) Images of sun formed on the ground by crevices in dense foliage of trees. (2) A single luminous point gives image of aperture.

14. *Camera obscura.*

15. *Intensity of Illumination.*—

a. The *intensity* of light is *directly proportional* to the *amplitude* of the luminous wave.

b. The *intensity* of light is *inversely proportional* to the *square* of the *distance* from the luminous body.

If the light of a candle [Fig. 45] passes through a square orifice illu-

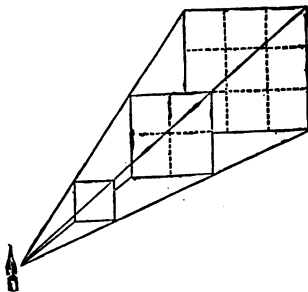


FIG. 45.

minated with the same quantity of light; at a single distance it will be equal to 1 square; at double distance, 4 squares; at triple distance, 9 squares. Hence the quantity of light is divided by 1, 4, 9, etc.

16. *Comparison of Intensity of Illumination.*—

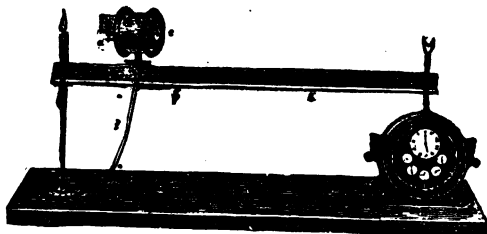


FIG. 46.

a. Method of shadows.

b. Method of oil spots; Bunsen's photometer [Fig. 46].

17. *Visual Angle*.—The angle contained between the two lines O A and O B [Fig. 47], which are drawn from the eye, O, to the limits of the object at A B, is called the *visual angle*.



FIG. 47.

a. It increases with the size of the object, and decreases with its distance.

18. *Optical Angle*.—The angle made by the lines drawn from the luminous point A [Fig. 48], to both eyes, B and C, is called the *optical angle*.

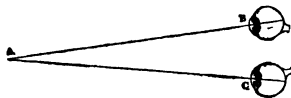


FIG. 48.

a. It diminishes with the distance.

19. *Estimation of Distance and Size*.—The distance and size of an object is estimated by comparison with known objects by means of the visual and optical angles.

20. *Disposition of Light*.—When a pencil of rays of light fall on any substance it is divided into more or less unequal parts, according to the nature and condition of the substance. It may be:

- a. Reflected.
- b. Transmitted.
- c. Absorbed.

21. *Absorption of Light*.—Even the most transparent bodies absorb some rays of light, while the opaque bodies absorb most, if not all, of them.

- a. Absorbed light rays are mostly converted into heat.
- b. Very thick glass or very thick strata of air absorb much light.
- c. Very thin opaque bodies may transmit some rays of light:
 - (1) Gold films transmit green, blue, or violet rays.
- d. Colored glass absorbs all rays but those of a special color.
- e. The difference in the disposition of light by various objects renders them distinctly visible by contrast, even if they are opaque.

22. *Reflection of Light (Katoptrics)*.—A pencil of light falling upon any body is at least partially, and sometimes totally reflected. The reflection is more or less perfect, according to the conditions of the reflecting surface.

a. Polished metals are excellent reflectors, so are glass, water, clouds, air strata, etc.

b. The incident ray, I B [Fig. 49], and the reflected ray, B R, make

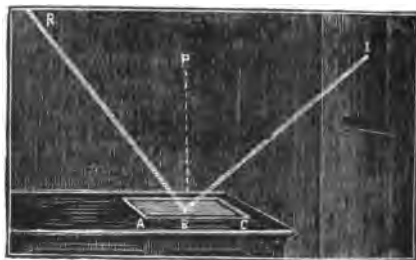


FIG. 49.

equal angles with a line perpendicular, as P B, to the reflecting surface, A B C, at the point of incidence, B. When the reflecting surface is curved, the line is perpendicular to the tangent at the point of incidence.

c. The incident and reflected rays and the perpendicular lie in the same plane.

d. On irregular surfaces the reflection is irregular and the rays are scattered.

e. Glass mirrors have double surfaces.

23. *Formation of Images by Reflection from Plane Mirrors.*—If a pencil of divergent rays A D [Fig. 50] be reflected at D, the rays D R

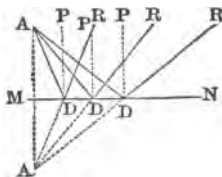


FIG. 50.

will appear to proceed from A an equal distance behind the mirror, M N, as the luminous point A is in front.

a. The *virtual image*, having no real existence, is formed by plane mirrors. Some of the rays entering the eye from all the many reflected rays, give a virtual image at the same apparent distance behind the mirror as the real object is in front of it; for the eye, being accustomed to straight direction of vision, receives the impression that the reflected rays proceed from points behind the reflector [Fig. 50].

b. Multiple images: When two plane mirrors are inclined towards each other with a luminous object between them, they produce several

images by reflection of the object by each mirror, and by reflecting the reflection of each other. Counting the real object as one, the eye sees as many objects and images as the angle of the inclination of the mirrors is contained in the 360° of the whole circle. The number of images is, therefore, for the angle of inclination of the mirror $= N^\circ$, or $= \frac{360^\circ}{N^\circ} - 1$. When the two mirrors are parallel, the number is infinite, and is only limited by the decreasing intensity of light by repeated reflection, $\frac{360}{0} - 1 = \infty$.

c. Kaleidoscope (Debusscope), (Brewster).

d. Sextant.

e. Reflection goniometer of Wollaston, of Babinet, etc.

f. Heliostat (Silbermann).

24. *Reflection by Curved Mirrors.*—The angles of incidence and reflection, made with a line perpendicular to the tangent at the point of incidence, are equal. In spherical curves the perpendiculars coincide, with the radii and their prolongations.

a. Lantern experiments with reflected light.

b. The curved surface of a convex mirror disperses rays of light. The foci, or coalescing points of the reflected rays, lie behind the curved surface, and are, therefore, *virtual foci*.

c. The images formed by a convex mirror (M N) [Fig. 51] are *virtual, erect, and smaller* (a b) than the object (A B).

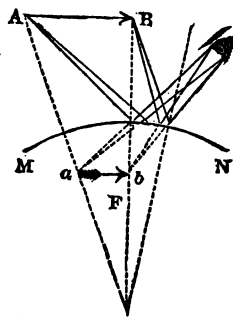


FIG. 51.

The size of any image formed by a spherical mirror, convex or concave, varies with the distance of the object. It is independent of the area. A large area increases the brightness.

d. The concave surface of a curved (spherical) mirror will reflect rays at equal angles with the radii.

If the luminous object is at an infinite distance, it will emit parallel rays, which are concentrated into a focus.

If the luminous body lies in the axis of the mirror, at C [Fig. 52],

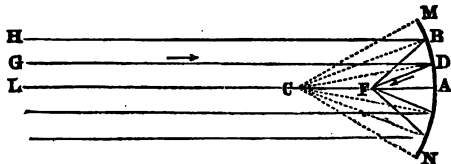


FIG. 52.

the principal focus will be half way between the center and the periphery, or at F.

When the luminous object is at a *finite distance* from concave mirrors, its divergent rays are rendered convergent, and are collected into a *real focus*. If the luminous point is beyond the center of curvature, to a focus between the center and the principal focus for parallel rays.

If the luminous point is between the center and principal focus, the reflected rays will be collected into a focus beyond the center. If the luminous point is at the center, the reflected rays will be returned to it.

If the luminous point occupies the *principal focus*, the reflected rays will be *parallel*, and there will be no collection to a focus.

If the luminous point is between the principal focus and the curved surface, the reflected rays will *diverge*, and there will be a *virtual focus* behind the mirror.

e. Images formed by concave mirrors are either real or virtual.

When the object emits parallel rays no image forms, but the light concentrates at the principal focus.

When the object is at a finite distance, *beyond center of curvature*, a *real image*, *inverted* and *smaller* than object, is formed between the center and the principal focus.

When the object is at the center, the *inverted* image is also at the center.

When the object is between the center and the principal focus, the image will be *real*, *inverted*, *enlarged*, and beyond the center.

No image is formed when the object is at the principal focus, the reflected rays being parallel.

When the object is between the principal focus and the mirror, a *virtual*, *erect*, *enlarged* image forms *behind* the mirror.

25. *Spherical Aberration*.—When the aperture of a mirror exceeds 10° , the reflected rays do not collect into one focus, but form a *caustic curve*.

Parabolic mirrors have no spherical aberration.

26. *Refraction of Light (Dioptrics).*—When the ray of light passes in an *oblique direction* from one transparent medium to another of *different density*, it is bent or refracted.

In passing from a *rarer* to a *denser* medium [Fig. 53], the ray is bent

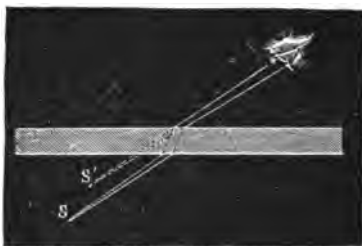


FIG. 53.

towards the line drawn through the point of incidence and perpendicular to the surface dividing the two media. When passing from a *denser* to a *rarer* medium, the ray is bent *from* the perpendicular.

Dense is understood as possessing more, and *rarer* as possessing less refractive power.

a. Law of sines—index of refraction: The angles made by the incident and the refracted ray [Fig. 54] with the perpendicular, have this

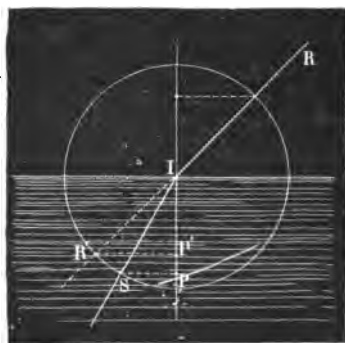


FIG. 54.

relation: For the same two media the *sines* of the angles have always the *same relative proportion to each other*, no matter what the obliquity. Hence, a perpendicular ray, having no sine, is not refracted. (Snellius, Holland, 1625.)

When a ray of light passes from air into glass the sines are, 3 : 2; from air to water, 4 : 3. This proportion, $\frac{3}{2}$ or $\frac{4}{3}$, is the *index of refraction* for these substances.

When reduced to comparison with a vacuum, the *absolute index of refraction* is found. It is also called *exponent of refraction*.

Table of absolute indexes of refraction.*

Vacuum	1.00000
Air	1.00029
Carbon dioxide	1.00045
Ice	1.30900
Water ($\frac{4}{3}$)	1.33600
Ether	1.35800
Alcohol	1.37400
Alum	1.45700
Benzol	1.50000
Canada balsam	1.53200
Crown glass ($\frac{3}{2}$)	1.53400
Quartz	1.54300
Oil of cassia	1.64100
Carbon disulphide	1.680 to 1.76800
Flint glass	1.664 to 1.83000
Diamond	2.43900
Lead chromate	2.97400
Cinnabar	3.20000

b. Total reflection:

Light rays passing from denser to rarer media are limited to a certain degree of obliquity. When this is exceeded, the rays can not pass out of the denser medium, but are totally reflected within.

The angle of obliquity is called the *limiting angle*.

Table of limiting angles.

From water to air	48° 35'
From alcohol to air	46° 52'
From benzol to air	41° 48'
From crown glass to air	40° 49'
From flint glass to air	37° 36'
From carbon disulphide to air	36° 31'
From diamond to air	23° 53'
From benzol to water	64° 23'
From carbon disulphide to water	52° 54'

c. Illustrations of reflection:

Glass tube with air immersed in water seems to be filled with mercury.

Rectangular glass prism.

Illuminated fountain.

Projection of arrow, distorted by refraction.

Coin in empty and filled vessel.

Estimation of depth of water [Fig. 55], shooting fish, etc.

d. Atmospheric refraction:

Real and apparent place of sun and stars;

Greater length of daylight.

Mirage—[Fig. 56].

Looming.

*All calculated for E Hne.

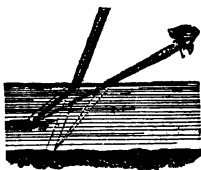


FIG. 55.

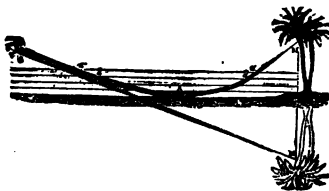


FIG. 56.

Refraction through glass, etc., bounded by *parallel planes*, and the consequent displacement of rays.

Refraction through glass, etc., bounded by *planes inclined at an angle*—*prismatic refraction*.

The ray in leaving the denser medium [Fig. 57] departs from the

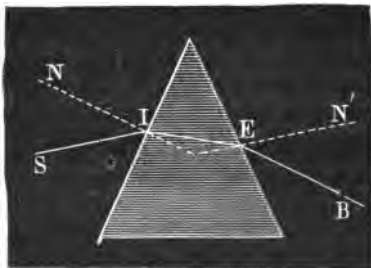


FIG. 57.

perpendicular direction, and as this is determined by the second surface, the direction is changed the more the greater the inclination of the two surfaces is to each other. The incident and refracted rays are always approaching the base, and receding from the refracting angle of the prism.

27. *Lenses*.—Lenses [Fig. 58] may be considered as compound



FIG. 58.

prisms. They may be of various forms, all reducible to prisms joined at the base or at the apex.

Convex lenses unite rays of parallel light to a focus, whose distance

varies according to the curvature and index of refraction. This is the principal focus.

The marginal rays do not all coincide—*spherical aberration*.

The best results of freedom from spherical aberration is when the two radii of a double convex lens have a length of 1:6. The *focus* is *real*.

Diverging rays are also collected into a focus, if the luminous point is further off than the principal focus. The luminous point and the point where the refracted rays converge are then situated at the *conjugate foci*.

When the luminous point is between the lens and its principal focus, the refracting rays will diverge and the focus will be *virtual*.

When the luminous point occupies the principal focus, the refracted rays will be parallel.

Real *images* are formed by *convex* lenses when the object is at a finite distance.

When the object is more than twice the distance of principal focus, it is *real*, *smaller*, and *inverted*.

When the object is between the principal focus and the lens [Fig. 59], its image is *erect*, *larger*, and *virtual*.

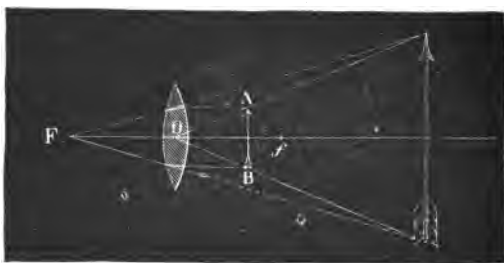


FIG. 59.

The size of an image is not affected by the area of the lens.

The *foci* of *concave* lenses are always *virtual*.

The *image* is always *virtual*, *erect*, and *smaller* than the object.

28. *Optical Instruments*.—The following optical instruments are based upon the principles which have been discussed: Camera obscura of photographers, camera lucida, microscope, telescope, magic lantern (stereopticon).

Galileo's telescope, 1610.

Solar microscope by Lieberkühn, 1738.

29. *The human eye; vision; stereoscopic vision; stereoscope; Wenham's binocular microscope; emmetropic, hypermetropic, and myopic eye*

[Fig. 60]; *corrections by lenses; accommodation by front curvature of lens being controlled by the ciliary muscles.*

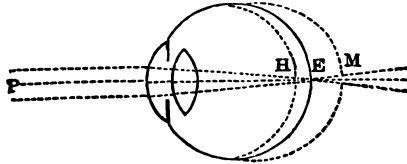


FIG. 60.

Deschanell's Natural Philosophy gives the following description of the human eye:

"The human eye [Fig. 61] is a nearly spherical ball, capable of turn-

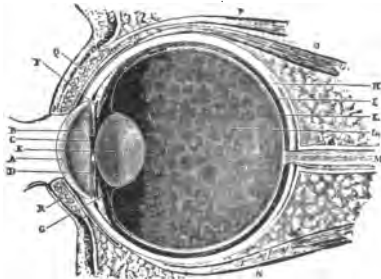


FIG. 61.

ing in any direction in its socket. Its outermost coat is thick and horny, and is opaque except in its anterior portion. Its opaque portion (H) is called the *sclerotica*, or in common language the white of the eye. Its transparent portion (A) is called the *cornea*, and has the shape of a very convex watch-glass. Behind the cornea is a diaphragm (D) of annular form, called the *iris*. It is colored and opaque, and the circular aperture (C) in its center is called the *pupil*. By the action of the involuntary muscles of the iris, this aperture is enlarged or contracted on exposure to darkness or light. The color of the iris is what is referred to when we speak of the color of a person's eyes. Behind the pupil is the *crystalline lens* (E) which has greater convexity at back than in front. It is built up of layers or shells, increasing in density inwards. This latter circumstance tends to diminish spherical aberration, the effect of which, in an ordinary lens, is to make rays which pass near the outside too convergent, as compared with those which pass near the axis. The cavity (B) between the cornea and the crystalline is called the anterior chamber, and is filled with a

watery liquid called the *aqueous humor*. The much larger cavity (L) behind the crystalline is called the posterior chamber, and is filled with a transparent jelly called the *vitreous humor*, inclosed in a very thin transparent membrane (the hyaloid membrane). The posterior chamber is inclosed, except in front, by the *choroid coat* or *uvea* (I) which is saturated with an intensely black and opaque mucus, called the *pigmentum nigrum*. The choroid is lined, except in its anterior portion, with another membrane (K) called the *retina*, which is traversed by a ramified system of nerve filaments diverging from the optic nerve (M). Light incident on the retina gives rise to the sensation of vision; and there is no other part of the eye which possesses this property."

Visual purple.

Impressions not instantaneous, nor ceasing at once.

Continuance of luminous impression about $\frac{1}{2}$ of a second.

Effects of the impressions: Lightning appears as a line; rockets, and other moving lights appear the same.

Toys based upon the *persistence* of vision: Thaumatrope, zœtrope, etc.

30. *Chromatic Dispersion by Prisms*.—The passage of a ray (S) [Fig. 62] of light through a triangular prism (P) not only displaces the

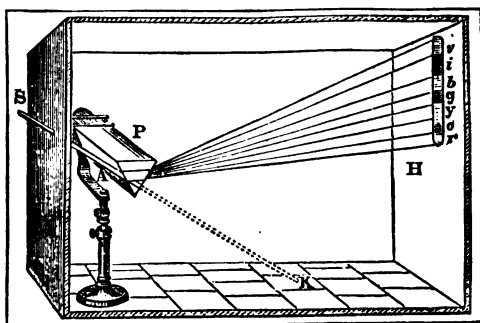


FIG. 62.

direction of the ray (*refraction*), but separates the white light into a number of color rays (H) (*dispersion*).

Of these rays *red* are the *least*, and the *violet* the most refracted from their original direction.

According to the older views of colors there were 7, or 3 principal colors.

According to recent views, each ray has a different wave length and velocity, producing a different color sensation.

There are many thousand different colors.

There are seven principal groups of colors: *Red, orange, yellow, green, blue, indigo, violet* [H, Fig. 62].

When a narrow slit is made to admit the light, colored images of the slit are produced; a wider slit allows overlappings and intermingling.

Simultaneous action of these colors gives the sensation of *white* light.

Illustrations of dispersion: Color spectrum prism [Fig. 62]; re-composition of white light by revolving color disc.

31. *Index of Dispersion*.—The separation of the extremes of color, as well as different intermediate portions, vary greatly by prisms of different media.

a. *Fraunhofer lines*.

b. Table of difference in indices of refraction and dispersion:

	Absolute refraction.*	Dispersion.
Vacuum	1.00000	
Air.....	1.00029	
Carbon dioxide.....	1.00045	
Ice.....	1.30900	
Water.....	1.33600	0.035
Alcohol.....	1.37400	0.029
Alum.....	1.45700	
Crown glass.....	1.53400	0.036
Quartz crystal.....	1.54800	0.026
Oil of cassia.....	1.64100	0.139
Carbon disulphide.....	1.76800	0.130
Flint glass.....	1.83000	0.052
Diamond.....	2.43900	0.038

32. *Chromatic Aberration in Lenses*.—The more refrangible violet rays [*v*, Fig. 63] are brought to a focus at *v*, nearer to the lens than the less refrangible red rays, *r*.

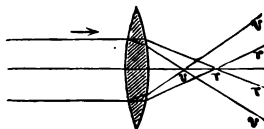


FIG. 63.

Hence, the double images with colored margins.

a. *Correction* is made by negative lenses of less curvature, but with greater dispersion, so that both foci coincide, and but a single image is produced, and it shows no color.



FIG. 64.

*E line.

These lenses are called *achromatic*. In Fig. 64, A is flint glass and B crown glass.

b. Natural phenomena due to dispersion; rainbow; halo.

33. *Colors of Natural Bodies*.—

a. Transparent media: Colored glass transmits only certain color rays, and absorbs the others.

b. Opaque colored media reflect certain color rays, and absorb the others.

Lantern experiment with colored ribbons.

34. *Complementary Colors*.—When all of the colors of the spectrum unite, white light is produced.

When a portion is missing, and the rest are united, intermediate tints are produced.

Any two portions of the spectral colors which, when united, will produce white light, are complementary to each other.

The experimental proof is best given by the double image prism and polarized light.

Red is complementary to green.

Orange is complementary to blue.

Yellow is complementary to violet.

Complementary color phantoms. Modification of appearance of colors by their surroundings.

Influence in perturbed vision and hallucinations, color phantoms, etc.

Harmony of color.

35. *Distribution of Rays in Spectrum*.—Only a portion of the spectrum [Fig. 65] is visible.

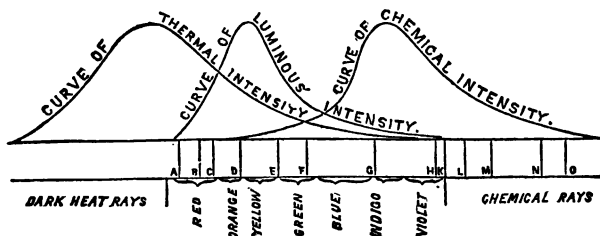


FIG. 65.

The *ultra red* rays are dark heat rays.

The *ultra violet* rays are actinic or chemical rays.

The *Fraunhofer lines* were mapped in 1814 (discovered by Walliston in 1802), and nine principal groups named.

Their importance was shown by Kirchhoff, in 1860.

a. Table of wave lengths and velocities of the principal color rays:

Color rays.	Number of waves in 1 inch.	Number of waves in 1 second of time.
Extreme red.....	38,640	442,000,000,000,000
Middle red.....	39,180	458,000,000,000,000
Orange.....	41,610	480,000,000,000,000
Yellow.....	44,000	517,000,000,000,000
Green.....	47,460	558,000,000,000,000
Blue.....	51,110	599,000,000,000,000
Indigo.....	54,070	634,000,000,000,000
Violet.....	57,490	675,000,000,000,000
Extreme violet.....	59,750	702,000,000,000,000

As light moves 185,000 miles (11,721,600,000 inches) in one second of time, and 63,360 inches make a mile, the number of waves in one second of time is easily calculated when the wave length is known.

b. Interference and measurement of wave lengths: When light passes through a narrow slit, the rays of light spread out from this as if it were a luminous point, but the oblique rays *interfere* so as to be behind each other for fractions of a wave length. Hence a spectrum of alternate light and dark spaces forms on the screen, *inflection*.

If white light be used, the spaces are colored with the different colors in succession; if monochromatic light be used, they are dark and of different distances from center for each color. Hence, the possibility of measuring wave lengths.

Newton's rings are caused by interference by reflection between a curved and a plane surface.

In thin plates of mica, in soap films, oil films on water, etc., the same phenomena occurs.

36. *Spectrum Analysis*.—Discovered by Kirchhoff and Bunsen in 1860, by experiments to explain phenomena previously observed.

White light from incandescent solids gives continuous spectrum.

Light from flames colored by metallic vapors gives interrupted spectra, some consisting of only one or a few bright lines.

a. Relation between dark Fraunhofer lines and bright metallic spectra:

Sodium, yellow, D.

Potassium, red, near A and B.

Hydrogen, C red, F greenish-blue, and near G.

Magnesium, δ 3 green.

b. Coincidence of solar and terrestrial spectra:

	Lines.
Iron.....	450
Barium.....	11
Titanium.....	118
Nickel.....	33
Calcium.....	75
Magnesium.....	4
Chromium.....	18
Hydrogen.....	4

c. Analysis by means of flame spectra: Each element has its own spectrum, which changes with the temperature, but does not coincide with that of another element, so that several spectra may be jointly present.

d. Absorption spectra of liquids or transparent solids: Peculiar bands are shaded out of color spectrum, by means of which many substances may be recognized. Among these substances are: Blood, anilin colors, permanganate of potassium, chlorophyl.

e. Applications of spectrum analysis: Spectra of sun, photosphere, chromosphere, corona, stars, comets, nebulae (carbon), solar storms (displacement of hydrogen F line), aurora borealis (yellow line), measurement of movement of fixed stars, rainbands.

Direct vision spectroscope [Fig. 66].



FIG. 66.

Curtman's Chemical Analysis says: "For ordinary purposes of analysis small direct vision spectroscopes are quite sufficient. Their price is so low as to bring them within the reach of every analyst. They consist of a cylindrical tube, with ocular (O) [Fig. 67], contain-



FIG. 67.

ing from 3 to 7 prisms (P), a draw tube with adjustable slit (S), and a collimator lens (C), placed between them to render the rays parallel."

Kirchhoff and Bunsen's spectroscope [Fig. 68].

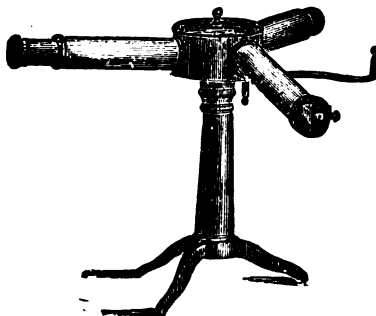


FIG. 68.

Microspectroscope.

Absorption spectroscope.

Projection spectroscope.

Star spectroscope.

f. Spectra produced by ruled gratings (*interference spectra*): Different degree of dispersion from refraction spectra; red end larger, blue end more condensed.

Diffraction spectra.

Irradiation is a subjective phenomenon due to arrangement of retina.

37. *Fluorescence*.—This is a property of some substances to convert the ultra violet invisible rays into visible ones of greater wave length and less velocity of vibration.

Table of wave lengths of ultra violet rays:

	Millimetres.	Inches.
H.....	= 0·000396	= 0·00001584
M.....	= 0·000366	= 0·00001464
N.....	= 0·000359	= 0·00001436
R.....	= 0·000309	= 0·00001236

Discovered by Stokes in fluorspar, hence the name.

Quinine shows blue fluorescence.

Uranium shows green fluorescence.

Chlorophyl shows red fluorescence.

Aesculin shows blue fluorescence.

Fluorspar shows blue fluorescence.

Electric light abounds in ultra violet rays of light of even greater refrangibility than those of sunlight. Hence, its great adaptibility to experiments with fluorescent substances.

38. *Phosphorescence*.—Some substances when illuminated for awhile (by solar or electric light) show, on withdrawal of light, a more or less prolonged luminosity, often of different color from the source of light.

The sulphides of barium, strontium, and calcium are especially phosphorescent.

Many other bodies show phosphorescence in Becquerel's phosphroscope.

Phosphorescence in rotten wood, insects, infusoria, etc., is a different phenomenon.

39. *Actinism—Photography*.—The invisible ultra violet rays, as well as the refrangible visible ones, decompose many chemicals—silver salts, mercurous salts, santonin, asphaltum, gelatin treated with chromates, etc.—and form the active agent in photography. Red or yellow glass does not permit their passage, and hence protects against their action.

40. *Calorescence*.—The invisible ultra red heat rays (calorescent

rays) may be concentrated by lenses and heat platinum, etc., to so high a temperature that visible light rays are emitted.

Tyndal's experiments.

41. *Double Refraction*.—(Discovered by Bartolin at Copenhagen in 1669.) Iceland spar [Fig. 69] and many other substances refract light



FIG. 69.

in two different pencils. One according to the ordinary laws of refraction, the *ordinary ray*; and the other not subject to them, the *extraordinary ray*, having a variable index of refraction.

a. Uniaxial crystals (one optical axis): Iceland spar, quartz, tourmaline, etc.

b. Biaxial crystals (two optical axes): Niter, sugar, strontianite, nickel salts, etc.

c. Double image prisms of Iceland spar: Rotation of one image around the other; use in Leeson's goniometer.

d. Construction of Nichol's prism, which passes only the extraordinary ray.

42. *Polarization of Light*.—Discovered by Malus, in 1808.

a. Polarization by reflection: In ordinary light the rays vibrate in planes intersecting each other at every angle [Fig. 70].

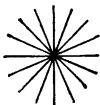


FIG. 70.

In polarized light the rays separate into two bundles, each having parallel rays [Fig. 71], the two directions of vibration being at right angles to each other.



FIG. 71.

When a ray of light falls on a plate of glass at an angle of $35^{\circ} 25'$, so that the reflected ($35^{\circ} 25'$) and the refracted ($54^{\circ} 35'$) ray make an angle of 90° to each other, both rays are polarized.

A bundle of thin plates renders polarization more complete than a single one, unless it is black (opaque). But every refracted ray, no matter at what angle, or by what substance, is partially polarized.

When polarized light is observed by the naked eye, it does not appear to differ from ordinary light. But when it again meets a substance capable of polarizing, either by reflection or by refraction (the analyzer), the peculiar properties are observed.

The polarized light will be reflected or transmitted only when the analyzer is in one position to the polarizer (parallel axes), and is extinguished in a position at right angles to this.

When films of doubly refracting substances are interposed between the polarizer and the analyzer, colors are produced by interference.

If thicker strata of certain organic solids or liquids are interposed, the analyzer has to be rotated several degrees to restore previous conditions (*circular polarization*) [Fig. 72].



FIG. 72.

Application to saccharimetry

Tourmaline plates.

Ring systems in crystals; calc-spar [Figs. 73 and 73a].



FIG. 73.



FIG. 73a.

Detection of reflected light in heavenly bodies.

Observation of signals, etc.

Examination of glass for telescope lenses, etc.

Tourmaline spectacles to prevent the glare of water or glass.

b. Circular polarization: Turning the polarized light to the right or left is manifested by

Quinine.....	— left.
Levulose.....	— left.
Galactose.....	— left.
French turpentine oil.....	— left.
Cinchonine.....	+ right.
Dextrose.....	+ right.
Sucrose.....	+ right.
Americanturpentine oil.....	+ right.

An electro-magnet turns the plane of polarization (Faraday).

XV. HEAT.

1. *Nature of Heat.*—Heat is a mode of motion affecting the *ether* surrounding the molecules and the molecules themselves.

The heat waves are *longer* and the vibrations *slower* than those of light.

The heat waves are invisible and less refrangible than those of red light.

Heat waves produce the sensation of warmth, but they are not perceived by the retina.

There were other hypotheses of heat: Caloric, material, imponderable.

2. *Sources of Heat.*—Solar heat is accompanied by light and chemical rays.

The earth receives only one part in 2,300,000,000 of the total heat radiated from the sun.

This amount of heat is sufficient to melt, every year, a crust of ice 95·6 feet (29·2 metres) thick over the surface of the earth (0·00728 inch every minute).

Pyrrheliometer.

Rosetti estimates the temperature of the sun at 10,000° C. (18,000° F.).

There are various theories of production and sustaining of the solar heat.

One is from *concussions by interplanetary matter* falling into the sun.

If all the planets were to fall into the sun, they would maintain the present heat for 46,000 years.

Another is the *combustion theory*.

If the sun consisted of coal, and this would burn with sufficient oxygen, it would maintain the present heat for 4,600 years.

Helmholz's *contraction theory*.

Two hundred and fifty feet shortening of the solar diameter per year would suffice to maintain the present heat.

Sun's diameter = 882,000.

Earth's diameter = 7,925 miles.

Sun's diameter = 112.79 times larger than that of the earth.

Earth's distance from sun = 91,718,000 miles.

Sun's revolution around its axis = 25.83 days.

3. *Terrestrial Heat*.—A thermometer placed in a deep cellar, or other sheltered excavation, some distance below the earth's surface, is not subject to variations of temperature from seasons.

The degree indicated will nearly coincide with the *mean annual temperature*.

Table of mean annual temperatures:

	Centigrade.	Fahrenheit.
Madras, India	27.8	82.1
Calcutta, India	25.5	78.0
Vera Cruz, Mexico	25.0	77.0
Rio Janeiro, Brazil	23.1	73.5
Canton, China	21.1	70.1
Tucson, Ariz.	20.5	69.0
New Orleans, La.	20.5	69.0
Austin, Tex.	19.5	67.0
Mobile, Ala.	19.0	66.0
Palermo, Sicily	17.2	63.4
Little Rock, Ark.	17.0	63.0
San Francisco, Cal.	12.8	55.0
St. Louis, Mo.	12.8	55.0
Baltimore, Md.	11.6	54.0
Paris, France	10.8	51.5
Indianapolis, Ind.	10.5	51.0
London, England	10.4	50.8
Des Moines, Ia.	9.5	49.0
Boston, Mass.	9.0	48.0
Sitka, Alaska	7.5	46.0
Madison, Wis.	7.5	46.0
Helena, Mont.	6.0	43.0
St. Petersburg, Russia	3.5	38.3
Irkutsk, Siberia	— 0.2	31.7
Iakutsk, "	— 9.7	14.6
Melville Islands, North America	— 18.7	— 1.66

Experiments made in Paris, Strassburg, Zurich, and Brussels show that there are no more variations of annual extremes at 24 metres deep.

The thermometer in the vaults of Paris observatory has shown 11.8° C. for half a century unchanged.

On going deeper, an average of 90 to 100 feet gives an increase 1° C.; 50 to 55 feet of 1° F.

If this increase continues proportionately, then, at the depth of 1 mile (5,280 feet), it would be 58.6° C., 137.5° F. (for 90 feet), or 52.8° C., 127° F. (for 100 feet); at 2 miles (10,560 feet), water boils; at 9 miles, iron is red-hot; at 24 miles, granite melts; at 40 miles, all known substances become fluid.

4. *Proofs of Increase of Internal Heat of Earth*.—Warm springs from greater depth.

Well at Neusalzweck (Minden), 2,219.5 feet, or 2,000 feet below ocean level = 33° C., 92° F., while the mean annual temperature = 9.5° C., 49° F.

Artesian well at Grenelle (Paris), 1,794·5 feet deep, water = 27·8° C., 82° F., while the air = 10·8° C., 51·5° F.

A geyser in Iceland = 100° C.

A spring in Padua = 99° C.

A spring in Karlsbad = 69° C.

Eselsschacht, Kuttentberg, 3,545 feet deep.

Well at Ho-Tsing, 3,000 feet deep.

Apendale coal mine, 2,175 feet deep.

Well at Insane Asylum, St. Louis, Mo., is 3,837 feet deep; temperature, 40·5° C., 105° F.; temperature of air at surface, 12·8° C., 55° F. (latitude, 38° 36' N.; longitude, 89° 36' W. Greenwich).

Table of irregularities in increase of temperature at Insane Asylum well:

English feet.	Fahrenheit.	Centigrade.
3,029.....	107·0	41·88
3,127.....	106·0	41·10
3,264.....	106·0	41·10
3,376.....	106·0	41·10
3,473.....	105·0	40·50
3,563.....	105·0	40·50
3,604.....	105·0	40·50
3,641.....	104·5	40·30
3,728.....	105·5	40·83
3,800.....	105·0	40·50
3,837.....	105·0	40·50

Table of increase of temperature at bore for rock salt at Spereberg, near Berlin:

Rhenish feet.	Centigrade.	Fahrenheit
700.....	19·500	67·000
900.....	22·310	73·000
1,100.....	24·830	76·800
1,300.....	27·420	81·000
1,500.....	29·787	85·000
1,700.....	32·066	89·600
1,900.....	34·140	93·400
2,100.....	36·130	97·000
3,390.....	45·945	114·600
4,042.....	48·125	118·625

This bore is 1,268·6 metres, 4,042 Rhenish feet, 4,162·7 English feet deep. The mean annual temperature is 8·6° C., 47·48° F.

The earth is flattened at the poles (*oblate spheroid*); is such as a liquid mass would produce at its rate of rotation. (Objection is made that erosion of solids might produce the same effect.)

Volcanoes used as proof of internal heat. (Various hypotheses.)

Earthquakes are explained by earth's liquid interior (also by undermining by solvent powers of water).

Elevation of mountain chains used as proof. (Objections made.)

5. *Mechanical Production of Heat*.—Heating of tools (saws, gimlets, etc.), of journals, of axles.

Indian mode of producing fire.

Tyndall's experiments on whirling table; boiling water in tube by friction.

Sliding down rope.

Fire syringe.

Meteors; falling bodies.

Friction matches.

Hammering.

Experiments of Count Rumford (Benj. Thompson) at Munich arsenal in 1798; boring brass cannon and heating water.

Sir Humphrey Davey melting ice by friction *in vacuo*.

Joule and Mayer's discoveries of mechanical equivalent of heat, in 1840 to 1843; increase of 1° F. of 1 pound of water equals mechanical force of 772 foot pounds, or 1° C. equals 423.65 kilogramme metres.

6. *Heat Produced by Chemical Affinity*.—Heat is produced by the coming together of atoms to form molecules. Heat is consumed in separating atoms or molecules from each other; hence, combustion is our principal source of artificial heat.

Illustrations of heat from chemical affinity: Phosphorus dissolved in carbon disulphide and poured on paper inflames; combustion of gas; compound blow-pipe; sugar and chlorate of potassium inflame on adding sulphuric acid; sulphuric acid and water produce heat; lime and water produce heat; potassium or sodium and water produce heat.

7. *Production of Heat by Conversion of Other Forces Into This Mode of Motion*.—Absorption of light.

Electric current heating bad conductors of heat, such as platinum wire; igniting gases; heating carbon points.

8. *General Remarks on the Correlation of Forces and Conversion of Energy*.—One kilogramme metre equals 7.233 English foot pounds.

9. *Effects of Heat*.—Heat causes greater energy of motion of the molecules, hence further separation in expansion of bodies.

Expansion of solids.

Linear expansion; cubic expansion.

Table of difference in rate of expansion when heated from 0° C. to 100° C.:

Zinc	$1/333$ to $1/240$	Lead	$1/351$
Tin	$1/462$ to $1/516$	Silver	$1/524$
Brass	$1/520$ to $1/535$	Copper	$1/584$
Gold	$1/645$ to $1/682$	Iron	$1/819$
Crown-glass	$1/1090$	Flint-glass	$1/1248$
Platinum	$1/1167$	Firwood	$1/2851$

—[LAVOISIER AND LAPLACE.]

Experiments illustrating expansion of solids by heat:

Ring and ball.

Compound bar of different metals.

Bologna flask and Prince Rupert's drops (glass tears) [Fig. 74].



FIG. 74.

Annealed and unannealed glass.

Balancing wheel of chronometer.

Gridiron pendulum for compensation.

Breguet's metallic thermometer; spiral of silver, gold, or platinum.

Repair of *Conservatoire des Arts et des Métiers*, by Molard.

Different pyrometers [Figs. 75 and 76]: Wedgewood's, Daniel's.



FIG. 75.



FIG. 76.

Unequal expansion of crystals.

The *co-efficient of expansion* is the fraction of increase of volume for 1° C.

Cubic expansion = $3 \times$ linear.

Expansion of Liquids.—Absolute expansion (independent of containing vessel) is very different in different liquids. In some liquids it is somewhat greater for equal number of degrees at higher than at lower temperature.

Table of apparent dilatation in glass, from 0° C. to 100° C.

Substance.	Cubical expansion.
Mercury	1 in 64
Hydrochloric acid, specific gravity 1.37.....	1 in 27
Water	1 in 22
Sulphuric acid, specific gravity 1.85.....	1 in 17
Ether.....	1 in 14
Olive oil.....	1 in 12
Alcohol.....	1 in 9
Nitric acid, specific gravity 1.40.....	1 in 9

Water is not regular in its expansion near the freezing point; greatest density at 4° C. = 39.2° F.

Melted bismuth and some alloys act in a similar manner.

Expansion of Gases: All gases and vapors (separated from their liquids) expand alike for equal increment of heat, unless close to their point of condensation. (Gay-Lussac.) For every 1° C. they increase $\frac{1}{273}$ of their volume, or $\frac{1}{491.4}$ for every 1° F.

The more condensable gases and vapors vary slightly from this rate.

Consequences of expansion of air by heat: Lighter specific gravity of the expanded gas, hence, ascending currents in chimneys, etc., and compensating currents; ventilation of rooms, mines, etc.; winds; land and sea breezes; trade winds; wind-mills, etc.

Steam engine:

Newcomen's atmospheric, 1705. Watt's double action, 1763. Papin, of Cassell, 1687; his first steamboat destroyed by a mob at Minden.

Water by boiling expands into 1,800 times its volume of steam, or about one cubic inch into one cubic foot.

Balloons with expanded air; Montgolfier, 1783, Pilatre de Rozier.

Illustrations of expansion of gases:

Steam engine, glass model.

Steam engine, metal model.

Candle bomb.

Air thermometer.

Papin's pot for digesting at high temperature (1687).

Pharoah's serpents.

Hero's engine.

10. *Measurement of Heat.*—

Uncertainty of personal impressions. Springs and cellars in summer and winter. Meeting in hall, the person from hot room feels cold, the one entering from street thinks the hall warm.

Experiments with three bowls of water: One person holds hand into the warm, another into the cold water, for some time, then both into the lukewarm, whose temperature appears different to each.

Expansion of liquids in suitable tubes with large bulb and narrow tube, used for measurement.

Thermometer [Figs. 77 and 78].

FIG. 77.

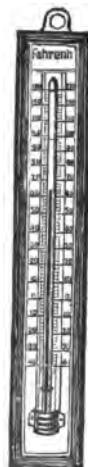


FIG. 78.

History: 1630, Cornelius Dubbel (Holland), nitric acid; 1694, Florentine Academy, Padua, alcohol; 1709, Fahrenheit (Danzig), mercury; 1730, Reaumur, alcohol; 1733, Delisle (St. Petersburg), inverted; 1742, Celseus (Upsala).

Fixed points: Freezing and boiling of water.

Air to be excluded from tube.

Change of bulb-tube influencing correctness.

Different scales and their equivalents and conversion:

$$F. = \frac{9}{5} C. + 32 = \frac{9}{4} R. + 32.$$

$$C. = \frac{5}{9} (F. - 32) = \frac{5}{4} R.$$

$$R. = \frac{4}{9} (F. - 32) = \frac{4}{5} C.$$

The fixed points are:

Fahrenheit, freezing water, 32° ; boiling water, 212° ; interval, 180° .

Reaumur, freezing water, 0° ; boiling water, 80° ; interval, 80° .

Centigrade, freezing water, 0° ; boiling water, 100° ; interval, 100° .

Hence, the size of degrees are: $80 R. = 100 C. = 180 F.$ ($4 R. = 5 C. = 9 F.$).

But the reading of the scale is dependent on the distance from freezing point, so that in Fahrenheit 32° must always be taken into consideration.

The desiderata in a thermometer scale are such size of degrees as to avoid fractions in ordinary work, and the zero point low enough to avoid negative degrees.

11. Absolute Temperature and the Law of Charles.—

All gases under constant pressure vary in volume directly as their absolute temperature.

Above the freezing point of water, gases expand $\frac{1}{273}$ of their volume for every 1° C. of increase; at -272° C. their volume would be $\frac{1}{273}$ of their volume at 0° C. Taking this volume as the unit and -273° C. as the zero of the scale, we would have the present 0° C. = 273° and the present 100° C. = 373° of the absolute temperature scale.

The number of degrees on the absolute temperature scale would indicate their volume.

For Fahrenheit degrees the calculation would be based upon $\frac{1}{491}$ of volume for 1° F., placing the absolute zero at -459° or 491° below zero.

Clinical, maximum, and minimum registering thermometers.

12. Transmission of Heat.—

a. Conduction (solids): From molecule to molecule.

b. Convection (liquids and gases): Change of place of molecules.

c. Radiation: Heat waves in ether through space.

Conduction is generally more active in proportion to density. Best in solids, very little in liquids and gases.

Table of difference in rate of conduction:

	Wiedemann and Franz.	Depretz.*
Silver	1,000	973.0
Copper	736	898.0
Gold	532	1,000.0
Brass	231	
Tin	145	303.5
Iron	119	374.0
Zinc		363.0
Lead	85	179.6
Platinum	84	381.0
German silver	63	
Bismuth	18	
Marble		23.6
Porcelain		12.2
Fire-clay		11.4
Glass		3.0

Wood is a bad conductor of heat; so are ashes, gypsum, woolen, and linen.

Hence, wooden handles for metal tools, glass handles for platinum wire.

Clothing a bad conductor of heat.*

Norwegian stove.

Fire-proof safes.

Ice-houses.

Water coolers.

Conduction of wire cloth is one of the causes of flame not penetrating through narrow meshes.

*Depretz is less correct than Wiedemann and Franz.

In crystals the conduction is unequal in different directions if the axes are unequal.

Wood conducts heat better along than across the fibers. Bark is a still worse conductor than wood.

In liquids and gases the conducting power is very low. Proven by inserting air thermometer under water, on top of which alcohol burns.

Gases confined between porous solids are the worst conductors.

Experiments with conduction of heat: Glass and copper rod held in flame; conductometer of various metals, having balls attached.

Convection.

In liquids and gases heat is propagated by the expanded, and, hence, specifically lighter, particles ascending and being replaced by the heavier ones.

Currents are thus set up which convey the heated particles. Hence, in water in a vessel heated from below, ascending hot particles and descending cold ones are noticed.

In viscous substances heat is communicated slowly.

Drafts in chimneys, winds, etc.

Radiation.

The radiation of heat is subject to the same laws as light.

Reflection by mirrors; plain and curved.

Archimedes' burning mirrors at Syracuse.

Cooking by reflected sun rays at Paris.

Refraction, as by burning lenses.

Dispersion of heat rays by prisms; calorescent rays.

Difference in character of heat rays of different wave length.

Reflection of heat by different bodies.

Out of 100 units received there are reflected by:

At angle of 45°.		Another authority.	
Silver.....	97	Polished gold.....	76
Gold.....	95	“ silver.....	62
Brass.....	93	“ brass.....	62
Platinum.....	83	Rough brass.....	51
Steel.....	82	Polished and varnished brass.....	41
Zinc.....	81	Looking glass.....	20
Iron.....	77	Blackened glass.....	12
Cast iron.....	74	“ metal.....	6

The rest of the light is absorbed.

The laws governing the direction of reflection, angle, etc., are identical with those of light.

Refraction and transmission through diathermanous bodies.

Heat rays passing from rarer to denser medium are subject to refraction like light.

The heat spectrum shows difference in properties of various rays of different wave lengths.

Table of number of rays out of 100 units, passing through various substances.

Heat emitted by:

Substance.	Lamp at white heat.	Red hot platinum.	Copper at 734° F., 390° C.	Copper at 212° F., 100° C.
Transparent rock salt	92	92	92	92
Fluorspar	78	69	42	33
Benzol	54	23	13	..
Iceland spar	39	28	6	..
Plate-glass	39	24	6	..
Dark-green tourmaline	18	16	3	..
Quartz	38	28	6	..
Gypsum	14	5	..	..
Alum	9	2	..	..
Sugar candy	8	1	..	..
Ice	6	0.5	..	..

The rest of the rays are absorbed.

Watery vapor and the exhalations of flowers absorb much heat.

Luminous heat rays enter glass-covered hot beds, etc.; the dark rays are retained.

Radiation from solid sources of heat is a correlative of absorption.

The best radiators are the best absorbers.

Table of relative radiation of solids (not diathermanous).

Lamp-black	100	Clean lead	19
Writing paper	98	Polished iron	12
Sealing-wax	95	“ tin	12
Crown glass	90	“ gold	12
Red lead	80	“ silver	12
Graphite	75	“ copper	12
Tarnished lead	49		

Table of absorption of heat by gases:

Air	1	Carbon dioxide	90
Oxygen	1	Nitrogen monoxide	355
Nitrogen	1	Sulphuretted hydrogen	396
Hydrogen	1	Marsh gas	403
Chlorine	39	Sulphur dioxide	710
Hydrochloric acid gas	62	Olefiant gas	970
Carbon monoxide	90	Ammonia	1,195

—[TYNDALL.

Various results and applications of radiation of heat:

Pipes in which heat is conveyed must be bright.

Radiators, stoves, etc., are best covered with graphite, etc.

Dew, hoar-frost, etc., by radiation.

Exchanges of heat rays by neighboring bodies; temperature established when exchanges are equal.

Difficulty of heating a room at first. Value of different modes of heating houses.

Clothing of various materials.

13. *Change of State of Aggregation.*—

Solid, liquid, gas, disassociation.

Change of state of aggregation due to heat and pressure.

Melting point of solids at constant temperature for equal conditions of same substance.

Substance.	Melting point.	
	Centigrade.	Fahrenheit.
Platinum.....	{ 2,534.0	{ 4,593.2
	{ 2,000.0	{ 3,632.0
	{ 1,800.0	{ 3,272.0
English wrought iron.....	{ 1,600.0	{ 2,912.0
French wrought iron.....	{ 1,500.0	{ 2,732.0
Pure iron.....	{ 1,800.0	{ 3,272.0
Steel.....	{ 1,300 to 1,400.0	{ 2,552.0
Gray cast-iron.....	{ 1,200.0	{ 2,192.0
White cast-iron.....	{ 1,050.0	{ 1,922.0
Gold.....	{ 1,250.0	{ 2,282.0
	{ 1,040.0	{ 1,904.0
	{ 1,000.0	{ 1,832.0
Silver.....	{ 957.0	{ 1,754.0
Copper.....	{ 1,150.0	{ 2,102.0
Aluminium.....	{ 700.0	{ 1,292.0
Antimony.....	{ 432.0	{ 809.6
Zinc.....	{ 360.0	{ 680.0
	{ 417.0	{ 800.6
Lead.....	{ 334.0	{ 633.2
Cadmium.....	{ 321.0	{ 609.8
Bismuth.....	{ 256.0	{ 492.8
	{ 264.0	{ 507.2
Tin.....	{ 230.0	{ 446.0
	{ 235.0	{ 451.0
	{ 109.0	{ 228.2
Sulphur.....	{ 111.0	{ 231.8
	{ 114.5	{ 238.1
Sodium.....	{ 95.6	{ 204.0
Potassium.....	{ 62.5	{ 144.5
Phosphorus.....	{ 44.3	{ 111.7
Ice.....	{ 0.0	{ 32.0
Bromine.....	{ -22.0	{ -7.6
	{ -9.0	{ 15.8
Mercury.....	{ -40.0	{ -40.0

Plastic state before fusion.

Difficulty of determining accurate point of melting.

Difference in melting point and consolidation (freezing) point.

Some substances pass directly from solid to gaseous state under ordinary pressure.

Melting point of alloys below the average of constituents.

Alloy of sodium and potassium is liquid at ordinary temperature, melts at -8°C .

Rose's metal (1 part each of bismuth, lead, and tin) melts at 93.9°C .

Wood's metal (1 part cadmium, 4 parts bismuth, 2 parts lead, 1 part tin) melts at 70°C .

Some substances expand on solidifying: Ice, cast-iron, bismuth, type metal, antimony, tin, zinc.

Boiling point.

Influence of surface of vessel.

Adhesion to sides.

Adding angular substances to promote easy boiling.

Influence of pressure.

Under equal conditions of atmospheric pressure, boiling occurs at a fixed point.

Higher pressure, raises boiling point; lower pressure, lowers it.

Culinary paradox.

Boiling under exhausted receiver.

Vacuum pans for evaporating syrups, etc.

Papin's pot, boiling under pressure.

Hypsometry.

Measuring height of mountain by boiling point of water,

At Quito, height = 2,808 metres = 9,212 feet; water boils at 90.1° C.

At St. Bernard's Hospice, height = 7,700 feet; water boils at 92° C.

Table of boiling points at reduced pressures:

Water boils, degrees Centigrade.	Pressure, millimetres, mercury barometer.	Water boils, degrees Centigrade.	Pressure, millimetres, mercury barometer.
0.....	4.60	50.....	91.98
5.....	6.53	60.....	148.70
10.....	9.17	70.....	233.09
15.....	12.70	80.....	354.64
20.....	17.39	85.....	433.04
25.....	23.55	90.....	525.45
30.....	31.55	95.....	633.77
40.....	54.91	100.....	760.00

Table of boiling point under pressure:

Water boils, degrees Centigrade.	Tension, atmospheres, 760 millimetres.	Alcohol boils, degrees Centigrade.	Tension, millimetres, mercury barometer.
100.....	1.000	78.....	760.00
121.....	2.025	100.....	1,697.55
134.....	3.008	155.....	6,259.19
144.....	4.000		
152.....	4.971		
159.....	5.966		
225.....	25.125		

Leidenfrost's drops; spheroidal state; prevention of surface contact, by vapor atmosphere.

Normal boiling point under one atmosphere pressure:

	Centigrade.		Centigrade.
Mercury.....	350.0	Bromine.....	63.0
Sulphuric acid.....	338.0	Chloroform.....	60.0
Phenol.....	181.0	Carbon disulphide.....	43.0
Acetic acid.....	118.0	Ethyllic ether.....	34.9
Formic acid.....	100.0	Sulphur dioxide.....	10.0
Water.....	100.0	Carbon ".....	78.0
Alcohol.....	78.4		

Latent heat (Black, 1763).

During the passage of bodies from the solid to the liquid, or from the liquid to the gaseous state of aggregation, heat disappears from observation by the thermometer.

When a solid begins to melt, the temperature remains constant until the whole is melted, then the temperature increases and remains

constant until all is vaporized, then the temperature of the vapor again increases with the further consumption of fuel.

The heat thus supplied by the fuel consumed, and which does not increase the reading of the thermometer, does inter-molecular work. It shifts the position of the molecules, and reappears on condensation.

"To convert a pound of ice at 32° F. into water at the same temperature, would require an amount of heat, competent, if applied as a mechanical force, to lift the same pound of ice to a height of 110,000 feet, 1 ton of ice nearly 50 feet, or between 40 and 50 tons of ice 1 foot.

"To convert a pound of water at 212° F. into a pound of steam at the same temperature, would require an amount of heat which would perform nearly seven times that labor."—[Tyndall's *Glaciers of the Alps*, page 830.

Distillation [Fig. 79].



FIG. 79.

One volume of water forms 1,689 volumes of steam.

One pound of water at 0° C. and 1 pound at 78·9° C. forms 2 pounds of water at 39° C.

One pound of ice at 0° C. and 1 pound of water at 78·9° C. forms 2 pounds of water at 0° C.

One pound of steam at 100° raises 6 pounds of water from 10° C. to 100° C., forming 7 pounds of water at 100° C.

One pound of steam at 100° C. melts 6¾ pounds of ice at 0° C., forming 7¾ pounds of water at 0° C.

Solution absorbs heat.

Melting absorbs heat.

Evaporation absorbs heat.

Condensation restores the heat again; hence, freezing mixtures, artificial production of cold and ice, etc.

Carre's ice machines.

Wollaston's cryophorus.

Formation of clouds.

Table of latent heat:

	Solid to liquid, Centigrade.	Liquid to vapor, Centigrade.
Water.....	79.25	Steam.....537
Sulphur.....	19.37	Alcohol.....214
Phosphorus.....	5.08	Ether.....90
Tin.....	14.25	Acetic acid.....102
Lead.....	5.37	
Zinc.....	28.13	
Mercury.....	2.83	
Bismuth.....	12.64	

Condensation of gases.

Under pressure or by refrigeration, or by both combined, all gases are coercible into liquids.

Pictet and Cailletet, in 1877, coerced the last of the permanent ones.

Critical point of temperature: Above a certain point of temperature, some gases are not condensable, even by greatest pressure.

This temperature is called the absolute boiling point.

Table of absolute boiling point:

Substance.	Temperature, Centigrade.	Tension, atmospheres.
Carbon dioxide.....	30.92	74.0
Ether.....	200.00	37 to 38.0
Carbon disulphide.....	275.00	78.0
Alcohol.....	259.00	119.0
Water.....	423.00	
Sulphur dioxide.....	157.00	79.0
Chlorine.....	148.00	
Chloroform.....	260.00	54.9
Oxygen.....	140.00	273.0

A number of gases may be easily liquefied at moderate pressure, others by cooling alone. On their evaporation great amounts of heat are consumed and artificial refrigeration procured.

Ammonia requires for liquefaction at 10° C., 4.4 atmospheres; at 20° C., 8.5 atmospheres; and at 40° C., 15.5 atmospheres.

Air compressed by 2 atmospheres is heated from 20° to 85° C.

Air compressed by 2 atmospheres, cooled to 30° C., and expanded equals — 25° C.

Specific heat (thermal capacity).

"The amount of work done in the shifting of the molecules of a body, when expressed in mechanical work, is perfectly enormous. The lifting of a heavy weight to the height of 1,000 feet may be as nothing compared to the shifting of the atoms of a body by an amount so small that our finest means of measurement hardly enables us to measure it.

"Different bodies give different amounts of trouble, if I may use the term, in shifting their atoms and putting them in new places. Iron gives more trouble than lead, and water gives far more trouble than either. * * *

"To raise a pound of iron a certain number of degrees in temperature

would, in fact, require more than three times the absolute quantity of heat which would be required to raise a pound of lead the same number of degrees. Conversely, if we place a pound of iron and a pound of lead, heated to the same temperature, in ice, we shall find that the quantity of ice melted by the iron, will be more than three times that melted by the lead. In fact the greater amount of molecular work invested in the iron now comes into play, the atoms again obey their own powerful forces, and an amount of heat corresponding to the energy of these forces is generated."—[Tyndall's *Glacier of the Alps*, page 330.*

One pound of water at 10° C., mixed with 1 pound of water at 30° C. results in a temperature of 20° C.

One pound of water at 10° C., mixed with 1 pound of iron filings at 30° C. produces a temperature of 12° C.; the water gains 2° and the iron filings lose 18°.

One pound of water at 30° C., mixed with 1 pound of iron filings at 10° C. produce a temperature of 28°; the water loses 2° and the iron gains 18°.

Water requires nine times the quantity of heat force to be raised to the same degree of temperature as iron, but thereby obtains the power of giving off again nine times as much heat in melting ice and in other operations.

Illustration of specific heat: Cylinders or globes of copper (95), iron (113), zinc (95), tin (54), antimony (50), lead (31), and bismuth (30) are heated to an equal temperature in an oil-bath and then placed on a cake of wax; the iron, copper, and zinc will melt through—the others will stick.

Table of specific heat:

Solids (Elements).

Substance.	Specific heat, water = 1.	Atomic weight, hydrogen = 1.	Specific heat × by atomic weight.
Lithium.....	0.9408	7.00	6.59
Sodium.....	0.2934	22.99	6.74
Aluminium.....	0.2143	27.01	5.80
Phosphorus (melted).....	0.2120	30.96	6.57
Phosphorus (yellow).....	0.1887	30.96	5.82
Melted sulphur.....	0.2026	31.98	6.48
Potassium.....	0.1696	39.02	6.50
Calcium.....	0.1700	39.99	6.80
Iron.....	0.1138	55.91	6.30
Copper.....	0.0952	63.17	6.00
Zinc.....	0.0956	64.21	6.20
Silver.....	0.0570	107.68	6.10
Tin.....	0.0562	117.70	6.60
Antimony.....	0.0508	119.96	6.10
Gold.....	0.0324	196.16	6.40
Mercury (solid).....	0.0319	199.71	5.40
Lead.....	0.0314	206.47	6.30
Bismuth.....	0.0308	207.52	6.30
Uranium.....	0.0277	238.48	6.50

*As this valuable book is out of print, the extract is given in full.

Gases.

Substance.	Volume.	Weight.
Air.....	0·2375	0·2375
Oxygen.....	0·2405	0·2175
Hydrogen.....	0·2359	3·4090
Ammonia.....	0·2996	0·5084

In different states of aggregation:

Substance.	Solid.	Liquid.	Gas.
Water.....	0·5050	1·0000	0·4805
Bromine.....	0·0843	0·1060	0·0555
Alcohol.....		0·5050	0·4534

The results vary with the temperature of observation: Some elements as bismuth, silicon, and carbon are conformable at high, not at low temperatures.

Effects of specific heat:

Large bodies of water exert an equalizing influence on the temperature, climate, and winds.

Transportation of aqueous vapor by clouds, etc.

Thawing out ice by hot water or hot metals.

14. *Polarization of Heat Rays.*—

Analogous to that of light rays, but not much studied.

Melloni's experiments with bundles of mica (reflecting) and lenses of salt.

15. *Interference of Heat Rays.*—

Not much studied.

Knoblauch's experiments: Interference by gratings on rock-salt, glass, etc

16. *Thermo-electric Apparatus.*—

Melloni's thermopile, etc., to be mentioned under electricity.

XVI. MAGNETISM.

Close connection between electric and magnetic phenomena.

Loadstone (lodestone) is a magnetic iron ore (ferroso-ferric oxide Fe_3O_4), crystallizing in octahedra.

The word *magnetism* comes from Magnesia, the name of a town in Asia Minor.

Loadstone attracts iron and steel, showing poles of greatest attraction.

The force is called *magnetism*.

Artificial magnets are made by rubbing iron or steel with natural magnets.

Soft iron becomes temporarily, hard iron and steel permanently, magnetic.

A magnetized steel bar shows distinct polarity at the ends.

Both ends attract non-magnetized iron, but the middle will not attract it.

When a magnetized steel bar is broken, either piece shows two poles.

Suspended magnets.

Magnetized bars, when suspended, will assume a north and south direction.

The end of such swinging magnet directed towards the north is called the *north pole*, and that directed towards the south the *south pole*, of the magnetic bar.

When two magnetized bars or needles are brought near each other, *equally named poles will repel, unequally named poles will attract* each other. Non-magnetized iron will be *attracted indiscriminately* by either pole.

Magnetic induction.

When an indifferent (non-magnetized) magnetic body, such as a piece of iron wire, is approached to a magnetized bar, the end nearest to the *north pole* will assume, by *induction*, *south-magnetic* polarity, while its opposite end becomes *north-magnetic* and will again induce magnetic polarity in another wire and attract it.

Hence, when a number of particles of iron filings are sifted over a glass plate or paper, and a magnet is placed below, each particle becomes polarized, and the result is a number of curved lines, showing the *curved lines of magnetic force*.

When two bar magnets with equal strength are placed together, so that the north pole of one covers the south pole of the other, no magnetic power will be exerted on outside objects, as each completely neutralizes the other. If one be stronger than the other, then the difference will be shown and affect direction.

Astatic needles for galvanometer.

To increase the attraction of magnets, the bar is curved into a horse-shoe form, so that both poles are brought nearer together and can act on the same piece of iron, called the keeper or anchor [Fig. 80].

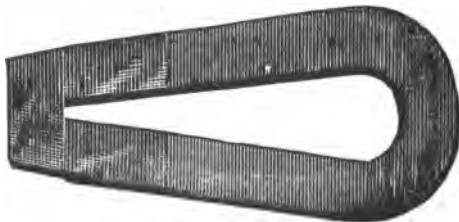


FIG. 80.

When a piece of iron is thus attached to each pole of a natural magnet, so as to have the poles act together, the attachment is called an *armature*.

A number of horseshoe magnets may be united into a *battery*.

Multiplying magnets in a battery does not proportionately increase the strength.

Table of Hæcker's Batteries.

Magnet weighing	2 ounces	bears.....	50 ounces.
"	8	"	140 "
"	16	"	200 "
"	4 pounds	"	31 pounds.
"	20	"	91 "
"	100	"	270 "

A large magnet constructed by Logemann for the Polytechnic School at Paris has 7 horseshoes joined, and weighs 67 kilos. It bears 275 kilos.

Tearing off keeper, leaving without keeper, jarring, and heating impairs magnets.

Substances attracted are called magnetic or *para-magnetic*, substances repelled are *dia-magnetic*.

A *para-magnetic* substance joins the poles, placing itself *axially*.

A *dia-magnetic* substance is repelled by both poles, and places itself equatorially, *i. e.*, at right angles, with the axis joining the poles.—[Faraday, 1845.

Para-magnetic.

Iron (steel, etc.)
Nickel.
Cobalt.
Manganese.
Chromium.
Cerium.
Titanium.
Palladium.
Platinum.
Osmium.
Aluminium.
Oxygen.
Paper.
Sealing wax.
Graphite.

Dia-magnetic.

Bismuth.
Antimony.
Zinc.
Tin.
Cadmium
Sodium.
Mercury.
Lead.
Silver.
Gold.
Copper.
Arsenic.
Phosphorus.
Iodine.
Chlorine.
Hydrogen.
Candle flame.

Terrestrial magnetism.

The earth is a great magnet, whose poles do not coincide with the geographical north and south poles.

Electric currents around the equator are supposed to be the cause.

The *north magnetic pole* was discovered by Sir James Ross, in June, 1831, on the west coast of Bæothia felix, north latitude, $70^{\circ} 5'$; longitude, $96^{\circ} 43'$ west of Greenwich, where Captain Ross observed a dip of $89^{\circ} 59'$.

The *south magnetic pole* was discovered by Ross in 1841 (not actu-

ally reached); south latitude, $75^{\circ} 30'$; longitude $15^{\circ} 40'$ east of Greenwich (South Victoria Land). Needle dip was $88^{\circ} 40'$.

The mariner's compass was known in Europe in the twelfth century; in China, 1100 B. C. (Fse Nah, indicator of the South).

The French call the pole pointing north the south pole of the needle.

Declination is the deviation of the needle from the true north pole, about 10° east at St. Louis. Periodic changes and diurnal variations.

Inclination or dip of needle: An equally balanced bar will dip when magnetized.

Dip at St. Louis, nearly 70° in 1860.

Dip at Rochester, N. Y., 75° .

Dip at City of Mexico, 50° .

If a balanced needle be magnetized, and then suspended so as to swing freely, it will both direct toward the north and dip.

The *intensity* of magnetic force is measured by the number of oscillations of the needle before coming to rest.

Magnetic induction of the earth.

Magnetizing of iron or steel bar, if placed in magnetic meridian and correct inclination. More readily when jarred, hammered, or filed. Railroads running north and south show magnetized rails.

Equilibrium in magnetized bodies free to move.

When a number of steel needles are equally magnetized; then pushed through corks and floated on the surface of water, there occur various positions, of more or less stability, when all the north poles of the floats are above water, and the south poles of a stronger magnet is held above them.

Connection of aurora borealis with terrestrial magnetism and frequency of sun spots.

Rays of aurora borealis occupying lines of magnetic force.

Aurora occurs together with high disturbances of magnetic needle. Magnetic storms.

Influence of heat on magnets.

At red heat iron loses its magnetism; at 20° to 25° C., manganese; at 350° C., nickel; at dark-red heat, chromium; at white heat, cobalt.

Influence of light.

Half of a steel needle placed in violet rays becomes north magnetic in some hours.

XVII. STATIC ELECTRICITY.

Amber (*ἡλεκτρον*) when rubbed shows attractive power. This property was known to the ancients; hence, the name electricity.

Various other substances act similarly. When rubbed they first

attract, then repel light bodies. Sealing wax, hard rubber, glass, and sulphur are examples.

In the electric pendulum, the pith ball suspended by a silk thread is attracted and then repelled by a glass rod which has been rubbed with silk, etc.

Opposite phenomena of glass and ebonite, sealing wax, or resin.

No electric attraction when metals are rubbed (unless insulated).

Conductors and non-conductors: When an electrified body is touched with a metal, the electricity is removed by conduction, while glass, resin, rubber, or silk does not remove it. Hence, division into conductors and insulators (isolators), or non-conductors. Discovered by Gray in 1729.

Conductors.

Silver	100.0	Charcoal	...
Copper	79.3	Graphite	...
Gold	65.0	Acids	...
Zinc	27.3	Water	...
Tin	17.0	Linen	...
Palladium	...	Cotton	...
Iron	12.0	All moist substances.	...
Lead	8.0	Aqueous vapor.	...
Platinum	8.0	The human body.	...
Mercury	2.0		

Insulators.

Hard rubber (ebonite).	Wax.
Shellac.	Glass.
Resin.	Silk.
Sulphur.	Fur.

Semi-Conductors.

Caoutchouc.	Dry paper.
Phosphorus.	Ether.
Dry wood.	Alcohol.

Electric polarity.

Discovered by Du Fay, in 1735.

Electricity, like magnetism, always appears in two modifications.

When glass is rubbed it becomes positively electrified, and the rubbing material negatively electrified.

When ebonite or resin is rubbed, it becomes negatively electrified.

Different hypotheses.

Franklin's unitary, + and —.

Du Fay's (1735) dualistic, vitreous, and resinous.

Bodies equally electrified *repel*, those *oppositely* electrified *attract* each other.

Illustrations of static electricity.

Bells [Figs. 81 and 82].

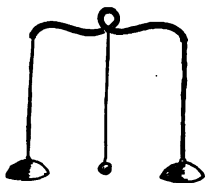


FIG. 81.

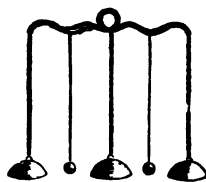


FIG. 82.

Electroscope.

Electric sec-saw and swing [Figs. 83 and 84].



FIG. 83.

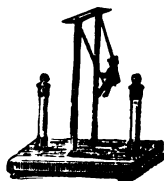


FIG. 84.

Dancing figures.

Hair repelled.

Insulating stool [Fig. 85].



FIG. 85.

Induction.

When an electrically excited body is approached to an indifferent one, opposite electricity is induced in the end nearest to the excited body, while the same electricity goes to the other end.

Electrophorus.

When the resinous cake is negatively electrified by fur, its metallic casing becomes positive. The cover, when put on, has + next to resin and — above. When it is connected with the casing the + and — unite. When taken off, free positive electricity resides in the cover and negative in the cake.

Electrical excitement is only on the surface.

Faraday's reversible bag.

Hollow globe, cylinder, etc.

Electrical intensity.

Electrical quantity.

Electrical potential.

Electrical machines [Fig. 86].

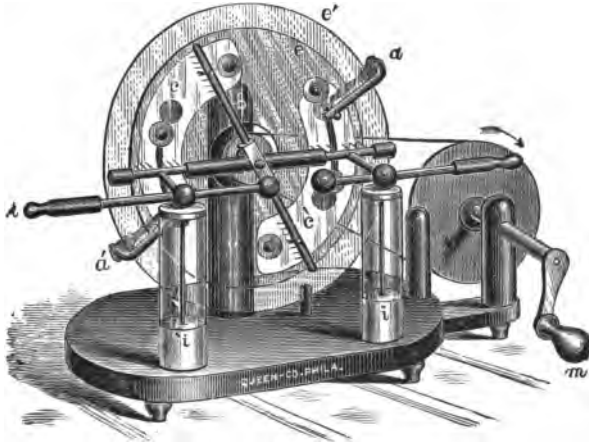


FIG. 86.

Very imperfect at first: Guericke's revolving sulphur ball, in 1671.

Wall's rosin cylinder, in 1708. Hawkesby's glass globe, in 1767.

Plate and cylinder machines.

Glass or vulcanite.

Rubbing cushion.

Prime conductor, secondary conductor.

A leather cushion covered with zinc, or tin amalgam, or tin sulphide (gold bronze) is used for rubbing the glass plates. The prime conductor is isolated. Rubbers connected with the second conductor and with the ground.

Wall, in 1708, first suggested that the electric spark is analogous to lightning.

Electric aurora.

Experiments with electric machine.

Electric gas pistol [Fig. 87].

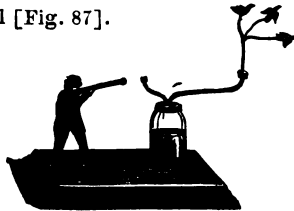


FIG. 87.

Igniting ether in spoon. Igniting illuminating gas (by man on insulated stool).

Spark to human body and other conductors.

Water spray.

Flame of candle blown out.

Electric whirl [Fig. 88].



FIG. 88.

Perforating card.

Points and rounded extremities.

Accumulators.

Leyden jars (Kleist's jars) [Fig. 89] were made by Kleist, of Kam-sin, Pommern, in 1745, and Cunæus, of Leyden (assisted by Professor Muschenbröck), in 1746.



FIG. 89.

Franklin's plate.

Dischargers [Figs. 90 and 91].

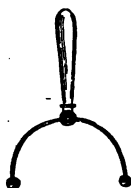


FIG. 90.

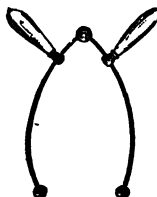


FIG. 91.

Experiments with static electricity.

Igniting powder.

Perforating card and glass plate.

Double burr (Lullius').

Lichtenberg's figures.

Directing rod.

Spark pictures by interrupted conductors.

Shape of spark. Path of least resistance.

Dissecting Leyden jars.

Electric spark.

Aurora tube.

Revolving electrophorus.

Holtz's influence machine.

Toepler's modification with metal buttons on plate.

Ozone formed by electric discharges.

Atmospheric electricity.

Thunder storms, etc.: Connections established by Dr. Benjamin Franklin, in June, 1752, by experiment near Philadelphia, Pa., with a kite. Similar experiments, by Franklin's advice, made in May, 1752, with an iron rod by French physicists. Richmann killed August 6, 1753, at St. Petersburg, in presence of painter Sokolow, by a similar experiment with a tall iron rod. At Franklin's suggestion, D'Alibard built a hut at Marly-la-ville, with an insulated iron rod 40 feet high. He observed sparks May 10, 1752. De Romas, at Nerac, used a kite with metallic string, June, 1753. The experiment was repeated in 1757 and produced a 9 to 10-foot spark. Romas was prostrated in spite of precautions taken, but survived the shock.

Clouds are generally + electrified, other clouds and the earth — by induction. Return stroke.

Lightning conductor: Conditions of safety; points above area of protection; material of rod; ground connection.

The area of protection has a base of a circle drawn with a radius twice as long as the point of the rod is above the base. Hence it forms a cone.

Velocity of electricity.

Differs according to intensity of charge. Highest observed, 288,000 miles per second.

Aurora borealis (australis): Supposed to be due to electric discharges in upper rarified air; particles arranged into rays coinciding with magnetic curves.

Supposed electric origin of cyclones, etc. Water-spouts.

In clear, cloudless weather the atmosphere contains + electricity. During fogs, the excitement is almost double. Precipitation of dew accompanied with electricity. Rain shows strong electricity, mostly + (71), sometimes — (69). Snow + (24), — (6).

Electric accumulation in human body.

Lighting gas jet after gliding about on insulating carpet, etc.

Electric phenomena in crystals by heating, pressure, etc. Tourmaline.

XVIII. DYNAMIC ELECTRICITY (GALVANISM, ETC.).

Ludovico Galvani, of Bologna, in 1789, made experiments with frog's thighs contracting by metallic contact (copper hook and iron railing). Experiments made by Alessandro Volta, of Padua, in 1799. Erroneous conceptions.

Chemical action produces electricity which has other phenomena than frictional (static) electricity, because it has a continual supply or current.

Galvanic cell or element.

Into a liquid [Fig. 93 or 94] two different substances are immersed,

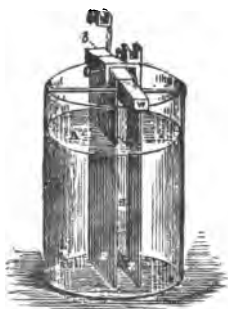


FIG. 93.



FIG. 94.

one of which is acted upon by the liquid more than the other. Liquid: Dilute sulphuric acid. Metals: Copper, not acted upon; zinc, dissolved. Current from zinc to copper inside of the liquid, returns from copper to zinc by connecting wire outside of liquid. Zinc is + in liquid; — in liquid (cathode). Copper is — in liquid; + outside (anode).

Poles are the ends of conducting wires. Zinc pole is —; copper pole is +. Electric current in double direction circuit. Closing and breaking the circuit.

Metallic plates equal electro-motors. Electro-motor series. (Differs

for different liquids; greatest tension between most distant in the following table.)

For sulphuric acid.	For hydrochloric acid.	For dichromate of potassium solution.
+	+	+
Zinc.	Zinc.	Zinc.
Lead.	Lead.	Copper.
Iron.	Iron.	Silver.
Nickel.	Copper.	Antimony.
Bismuth.	Bismuth.	Lead.
Antimony.	Nickel.	Bismuth.
Copper.	Silver.	Nickel.
Silver.	Antimony.	Iron.

Complete schedule of tension extending over many elements, but must vary according to circumstances.

—	P.	Sn.	Be.
O.	As.	Pb.	Mg.
S.	C.	Co.	Ca.
Se.	Au.	Ni.	Ba.
N.	Pt.	Fe.	Li.
Cl.	Hg.	Zn.	Na.
Br.	Bi.	H.	K.
I.	Sb.	Mn.	Rb.
F.	Cu.	Al.	+

The greater the chemical action the more *energetic* the current. The larger the electro-motors the greater the *quantity* of current. The greater the resistance to polarization and discharge, the more *intense* will be the current.

Conduction of galvanic current:

Silver.....	100·0000000
Copper.....	99·9000000
Zinc.....	29·0000000
Platinum.....	18·0000000
Iron.....	16·8000000
Lead.....	8·3000000
Carbon.....	0·0400000
Mercury.....	1·6000000
Dilute sulphuric acid.....	0·0009907
Strong nitric acid.....	0·0008868
Saturated solution of chloride of sodium.....	0·0003152
Distilled water.....	0·0000001

Galvanic batteries.

When two or more galvanic cells are united a galvanic battery is formed [Fig. 95].

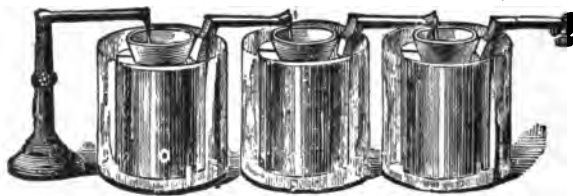


FIG. 95.

There is either an increase in *quantity*, if all equal plates are united, or an increase in *intensity* if alternate plates are united.

Principal forms of batteries: Single and double fluids.

Metals: Zinc and copper (Daniel), zinc and carbon (Bunsen), zinc and platinum (Grove), zinc and platinized silver (Smee); dipping with both plates, or zinc only dipping, etc.

Liquids: Dilute sulphuric acid; dilute sulphuric acid and chromate of potassium; sulphuric and nitric acids, etc.

Oxidizers used to prevent polarization of hydrogen.

Keeping batteries in order: Cleanliness. Metallic contact. Amalgamation of zinc. Care of carbon plates.

Ohm's laws (Berlin, 1827).

1. The quantity of electricity passing through a closed circuit (the strength of the current) is directly proportioned to the electro-motoric power of the battery (the electro-tension within the liquid).

2. The electro-motoric power (or strength of current) in the open circuit is proportional to the tension of free electricity at the poles.

3. The strength of the current is inversely proportional to the whole resistance in the liquid and in the conducting wires. Hence, strength

equals electro-motoric power divided by liquid resistance ($S = \frac{E}{L}$).

When a short circuit is used so that the resistance in the wire is small in proportion to that in the battery-jar, the addition of more alternating cups is of no advantage, while more are needed if the circuit becomes longer.

In a short circuit the increase of surface of elements increases the current.

Measurement of electric current.

By *voltameter*, measuring volume of gas from decomposed water.

By *rheometer*, measuring deflection of magnetic needle surrounded by current.

Rheostat is an apparatus for keeping equal currents by interpolation of different lengths of wire.

Units of Committee of British Association.

Ohm = unit of resistance = resistance of column of mercury 1 millimetre square, 1 metre long = 10,000,000 metre per second.

Volt = unit of electro-motoric power = somewhat less than a Daniel's cell.

Ampere = 1 volt passing through 1 ohm of resistance.

Weber = quantity of electricity conveyed by current due to a force of 1 volt through resistance of 1 ohm.

Farad = capacity of a condenser holding 1 weber, when charged with potential of 1 volt. (Too large for most purposes.)

Microfarad = $\frac{1}{1000}$ farad.

One mile of submarine cable has resistance of four to twelve ohms.

Heat Produced by Galvanic Current.—Wires conducting the current are heated; the more so the more resistance they offer. Hence, iron or platinum will glow with red or white heat, so as to ignite powder, etc.

Carbon points become incandescent, forming arc.

Platinum loops for galvanic cautery.

Electric light. Arc light. Platinum light. Carbon thread (Edison's) *in vacuo*.

Automatic regulators (Dubosq).

Electric current no longer produced by batteries, but by dynamo-electric machine [Fig. 96].

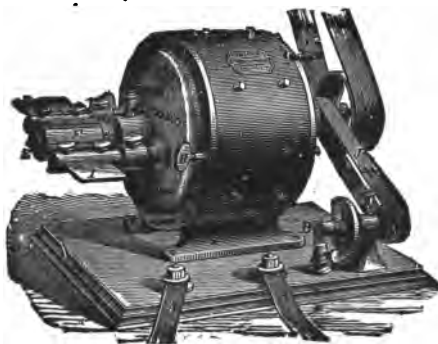


FIG. 96.

Spark at opening and closing of circuit.

Chemical decomposition by electric current.

Pure water not easily decomposed; acid should be added.

Water decomposed into 2 volumes hydrogen and 1 volume oxygen; ammonia into 3 volumes of hydrogen and 1 of nitrogen; hydrochloric acid into 1 volume each of hydrogen and chlorine.

Salts decomposed: Iodide of potassium into potassium and iodine. Iodine shown by starch solution; potassium decomposes water and

forms hydrogen and caustic potash; sulphite of sodium decomposed into sodium (which decomposes water into caustic soda) and sulphur tetroxide which becomes sulphuric acid and oxygen with water, shown by litmus solution or phenolphthalein.

Discovery of alkali metals by Sir Humphrey Davy, in 1807.

In metallic salts the metals are set free at zinc pole, the acid radicle at carbon pole. Application for electroplating with gold, silver, nickel, etc.

Galvanoplastic reproduction of metallic surfaces. Electrotyping.

Physiological effects.

Contraction of muscles when motor nerve is pervaded by current. Electrotonics.

Muscular electricity either assisted by like current or overcome by opposite.

Bipolar and unipolar application. Different effects of anode and cathode.

Electrolytic effect on living tissue.

Electro-cautery.

Different effects of constant and interrupted currents.

Effect on sensory nerves: Pain, taste, vision, noises.

Electric fishes: Raya (torpedo fish). Electric eel.

Electric shock, etc.

XIX. ELECTRO-MAGNETISM.

Oersted, 1820 (Copenhagen), discovered that a magnetic needle is deflected by passage of electric current.

Ampere's law:

A human figure through which current enters at the feet and passes out at the head, faces the needle. On passing the current the south pole of the needle is deflected to the right.

In passing current in one direction above needle and in an opposite direction below it, by a continuous wire, the deflection will be doubled, and by coiling the wire several times it will be multiplied in accordance with the number of coils, so that feeble currents may be detected.

The instrument is called a *galvanometer*. A static needle is used to make it sensitive.

A movable coil of insulated wire through which a current passes places itself into the magnetic meridian (De la Rive's Ring, Ampère's Solenoid). Two solenoids will attract or repel each other the same as two magnets.

The turns of a coil attract each other and shorten the length. Use for current interrupter.

Rotating apparatus by alternate repulsion and attraction between helix and permanent magnet, by reversal of current or of poles.

Magnetism induced in iron bar surrounded by helix in which current passes. Electro-magnetism (Arago, Sturgeon).

When perfectly pure soft iron is used in electro-magnet the attraction ceases at once on breaking the current, but returns at once on connecting. A piece of hard iron or steel becomes a permanent magnet.

If the iron lies loosely in coil it is drawn to central position when displaced.

Horseshoe electro-magnet obtains co-operation of both poles.

Various motor engines based on electro-magnetism.

Electro-magnetic telegraph [Fig. 97, key and sounder combined].



FIG. 97.

The electro-magnet at end of line alternately attracts and lets fall its keeper upon sending the current or interrupting it by a key (signal key).

History of electro-magnet: In 1808, Sömmering and Coxe, electrolysis; 1820, Oversted's deflecting needle (Laplace, Ampère, Ritchie); 1832, Baron Schilling's telegraph model exhibited before Emperor of Russia; 1833, Gauss and Weber, at Göttingen, 10,000 feet between physical cabinet and astronomical observatory; 1833 to 1837, Steinhell, in Munich, used ground connection instead of return wire, 40,000 feet; 1834, Wheatstone measured velocity by rotating mirrors, in 1835 he used Schilling's model in his lecture at King's College; 1836, telegraph model used in Munke's lectures at Heidelberg; 1837, Cooke built telegraph in England; 1832 to 1836, Morse made experiments with electrolytic telegraphing; 1836, Morse adopted electro-magnetic telegraph with alphabet of dots and dashes; 1850, first *submarine cable* laid by Brett, between Dover and Calais; 1868 (August 5), Atlantic cable completed.

Morse's alphabet consists of dots and dashes produced by short and long contacts of electro-magnet key pressing a point on a moving slip of paper.

Morse's register. Reading by sound. Relay and local battery. Dial telegraph. Printing telegraph (Baine).

Electric fire alarms, call bells, etc. Regulation of clocks; recording time by chronograph. Ball forceps with electric indicator of metal.

For long wires, alternating cup batteries must be used.

Varieties of telegraphing: Duplex and quadruplex. Tuning forks used.

Galvanic induction (Voltaic induction, Faraday, 1831).

A closed coil of insulated wire conducting a *primary* current will induce a *secondary* current in another helix of insulated wire surrounding it, for a short time and then discontinues. The secondary current flows in opposite direction to the primary. On breaking primary an *extra* current is induced by secondary. Sparks are induced at breaking.

Such induction spirals are much used for interrupted current (faradaic) for medical purposes.

Automatic magnetic hammer break-piece (Wagner, De la Rive).

Ruhmkorff's induction apparatus [Fig. 98] has insertion of iron rods

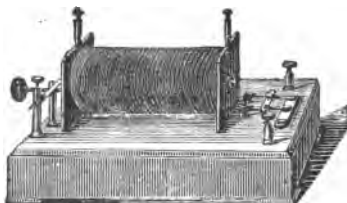


FIG. 98.

into primary helix of thick insulated wire. The secondary helix is of many coils of very thin wire. Automatic break-piece and condenser. Produces sparks at high tension. Used for firing fuses in mines, volatilizing metals, electric light *in vacuo*, igniting gas jets, etc.

Induction spark *in vacuo*: Geissler's tubes [Figs. 99 and 100].

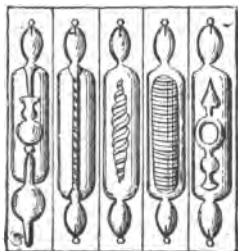


FIG. 99.

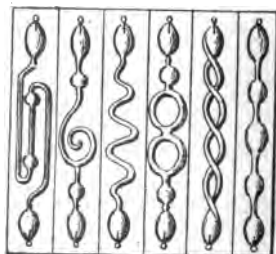


FIG. 100.

Aurora tube: Analogy of luminous appearance to aurora borealis.

Magneto-electric induction (Faraday, 1831).

When a magnetic bar is approached to an insulated helix electric currents are induced; also when a helix rotates before a permanent magnet so as to cut across the magnetic lines of force.

Various magneto-electric machines used for medicinal purposes.

Similar machines, called *dynamo-electric* [Fig. 96, page 79], now used for production of electric currents for electroplating and electric light. They are driven by steam and electro-magnets rotate between helices, producing enormous currents.

Storage batteries.

Faure, etc., produce in lead plates exposed to electric currents oxidation and reduction, so that a current is again produced by them.

XX. THERMO-ELECTRICITY.

In 1823 Seebeck, of Berlin, discovered that when a bismuth bar was soldered to a bent copper bar at both ends, the heating of one soldered end would produce an electric current and deflect a magnetic needle.

Table of thermo-electric tension of metals.

When heated at the soldered joint the current passes from the lower to the upper metal in following series:

Seebeck.	Becquerel.	Hankel.
+ Antimony.	+ Antimony.	+ Antimony.
Iron.	Iron.	Iron.
Zinc.	Zinc.	Silver.
Silver.	Silver.	Zinc.
Gold.	Gold.	Lead.
Tin.	Copper.	Tin.
Lead.	Lead.	Copper.
Mercury.	Platinum.	Gold.
Copper.	— Bismuth.	Platinum.
Platinum.		Mercury.
— Bismuth.		— Bismuth.

Nobili's thermo battery has bismuth and antimony rods soldered at alternate ends, and prevented from contact at all other points. It forms a compact battery for experiments with heat. Funnel guards are attached to ends and the poles connect with a galvanometer.

Marcus' thermo battery uses two alloys: Copper 65 and zinc 8 parts; antimony 12 and zinc 5 parts. Other batteries use iron and platinum or iron and silver.

Clamond's cylindrical battery has a lamp to heat the inner part.

Various instruments depending on electricity and magnetism.

The *telephone* has weak electric currents produced by vibration of plate before a magnet, reproducing the sonorous waves of the human voice at distant receiving station.

The *phonograph* registers the complicated waves of the voice on a cylinder coated with tin-foil, and reproduces them again through a telephonic mouth-piece.

The *photophone* is a telephone using a beam of light in place of connecting wire. Variations in light are produced by vibrating screen. Variable electric conduction of selenium under influence of varying light produces sound vibrations at receiving station.

XXI. DISSOCIATION.

As the temperature rises the molecules of *solids* are separated from each other (and most likely split up into smaller molecules), until the *liquid* state is reached. Then, again, with rise of temperature the molecules separate farther and are reduced to smaller aggregates and the *gaseous* or *vaporous* state is obtained. On further heating, still smaller molecules are formed, the atomic bonds are gradually loosened, and at first saturated, then non-saturated molecules are liberated into groups having a free atomicity, which at lower temperature unites them again into larger groups (condensed molecules). At still higher temperatures the atomic bonds are incapable of holding the atoms together and free atoms are formed, which can not unite till the temperature is lowered again.

It is suspected, and supported by Huggin's spectroscopic observations, that many of our elements at very high heat separate into simpler units, of which our present atoms are aggregates.

As an example we may take some nitrogen and chlorine compounds with oxygen, forming at low temperatures complex molecules, which, by a slight rise in temperature above their boiling point, separate (*dissociate*) into simple compounds with a free atomicity:

N_2O_3 and NO , N_2O_4 and NO_2 ;

Cl_2O_4 and ClO_2 ;

Hg_2Cl_2 and $\text{HgCl}_2 + \text{Hg}$;

H_2O and $\text{H}_2 + \text{O}$.

Iodine up to $600^\circ \text{C.} = \text{I}_2$; above $1,400^\circ \text{C.} = \text{I}$.

Sulphur vapor up to $500^\circ \text{C.} = \text{S}_8$, specific gravity, 6.65 (air = 1); at $1,000^\circ \text{C.} = \text{S}_2$, specific gravity, 2.216 (air = 1).

CHEMISTRY.

I. INTRODUCTION.

Chemistry is the study of the phenomena of the force of affinity. Other forces, such as mechanical motion, affect matter in masses, or the *molecules* composing the matter, their distances, relative positions, velocity of motion, attractions, and repulsions. Chemistry is concerned with the *atoms* composing the molecules. It studies the properties of atoms, their associations into molecules, their mutual attractions, the changes of their positions by application of forces, the compounds they form when brought together with other atoms or molecules, etc.

Definition of terms.

An *element* is simple matter—matter composed of atoms of a single kind, incapable of being subdivided into particles of different nature. Iron, sulphur, etc., are examples.

An *atom* is the smallest particle of *simple matter* capable of combining with another particle of simple matter. It is the *chemical unit*.

A *molecule* is the smallest particle of matter, *simple* or *compound*, capable of existing by itself in the free state. It is the *physical unit*.

Mechanical mixtures are produced by mingling together particles of different substances, without the union of their atoms into molecules.

Compounds are produced by uniting the atoms of different substances into molecules. Hence, matter may be *simple* or *compound*. We now (1888) know of about 73 elements or simple matters, while some others are claimed whose nature as yet is undecided, so that more may be added hereafter.

Each atom is of a definite weight, and most atoms differ from each other in weight. This is called the *atomic weight*. We can not weigh a single atom, but can compare equal numbers of atoms of one element with those of another. *Hydrogen* being the lightest is used as the *standard*. Atoms usually unite in small numbers to form molecules. The weight of a molecule of simple or compound matter is called the *molecular weight*.

Explanation of multiple proportions.

Law of Avogadro, 1811 (Ampere, 1814): In the *gaseous* state of matter equal *volumes*, under equal conditions of *pressure* and *temperature*, contain equal number of *molecules*.

The number of molecules in 1 cubic inch at 30 inches barometric pressure and 0° C. = 10^{23} = 100,000,000,000,000,000,000. The

size of molecules, according to Sir William Thompson, lies between $\frac{1}{250000000}$ and $\frac{1}{500000000}$ inch. Molecules, even in densest solid, are far apart in proportion to their size. Their interstices are filled with luminiferous ether, which forms an envelope for each. In gases molecules always occupy two volumes, no matter how many atoms they contain: $H + H = H_2$, or 1 volume of hydrogen + 1 volume of hydrogen = 2 volumes of hydrogen; 1 volume O + 2 volumes H = 2 volumes H_2O ; 1 volume N + 3 volumes H = 2 volumes H_3N . Hence, the molecular weight determines the *density* or specific gravity of gases.

Molecules are always in motion.

States of aggregation.

Size of liquid and of solid molecules most likely larger than that of gases.

Aggregation of several molecules to *condensation molecules*.

But few elements exist in nature in their free state. Most occur in compounds. New discoveries occur from time to time. Scarcity of material a bar to successful investigation of some. Recent discoveries by spectroscopy; Cs, Rb, Tl, In, Ga, etc.

Mode of determining atomic weight: (1) Density of vapor. (2) Equivalent proportion in compounds. (3) Specific heat of solids. (4) Isomorphism.

Table of chemical elements.

Stas' atomic weights, calculated with extreme accuracy (O = 16), are: Br = 79.75212; Cl = 35.36836; I = 126.5329; K = 39.03916; Li = 7.0014445; Na = 22.9854; Ag = 107.6602; N = 14.00889.

No.	Name.	Symbol.	Atomic weight.	Specific gravity.	Melts at $^{\circ}C$.	Boils at $^{\circ}C$.	Discoverer.	Time.
1	Aluminum	Al	27.01	2.56	700		Woehler.	1827
2	Antimony	Sb	119.96	6.71	437	1,100	Basilius Valentinus.*	1450
3	Arsenic	As	74.92	5.73	500	500	Schroeder.	1694
4	Barium	Ba	136.76	4.50	475		Davy.	1808
5	Beryllium	Be	9.09	2.10	960		Woehler and Bussy.	1828
6	Bismuth	Bi	207.52	9.82	264		Known since	1400
7	Boron	B	10.84	2.68	Very high.		Davy; Gay-Lussac and Thénard.	1807
8	Bromine	Br	79.768	3.18	-7.5	63	Balard.	1826
9	Cadmium	Cd	112.77	8.7	321	772	Stromeyer; Hermann.	1818
10	Cæsium	Cs	132.60	1.89	27		Bunsen and Kirchhoff.	1860
11	Calcium	Ca	39.99	1.577	500?		Davy.	1807
12	Carbon	C	11.97	3.52			Known to ancients,	
13	Cerium	Ce	140.42	6.92	1,000?		Klaproth.	1803
14	Chlorine	Cl	35.37	35.37†	-7.5	-33	Scheele.	1774
15	Chromium	Cr	52.01	6.8	2,000		Vauquelin.	1797
16	Cobalt	Co	58.88	8.5	1,800		Brandt.	1733
17	Copper	Cu	63.17	8.8	1,150		Known to ancients.	
18	Davyum?	Da	160?	9.39	?		Sergius Kern.	1877
19	Decipium?	Dp	171?				Delafontaine.	1877

* Described by Basilius Valentinus.

† The specific gravity of gases is given in heavy type, H = 1.

No.	Name.	Symbol.	Atomic weight.	Specific gravity.	Melts at °C.	Boils at °C.	Discoverer.	Time.
20	Didymium....	Di	146.8	6.54	1,000?		Mosander.	1839
21	Erbium.....	Er	165.89		?		Berzelius.	
22	Fluorine.....	F	18.984	18.984		—170	Schwankhard.*	1670
23	Gallium.....	Ga	68.85	5.9	30		L. de Boisbaudran.	1875
24	Germanium....	Ge	163?				C. Winkler.	1886
25	Gold.....	Au	196.16	19.32	1,040	?	Known to ancients.	
26	Hydrogen.....	H	1.0	1.00		—215	Paracelsus.	1500
27	Holmium?....	Ho	162.0	?		?	Cleve.	1879
28	Indium.....	Id					Websky.	1884
29	Indium.....	In	113.4	7.4	176		Reich and Richter.	1863
30	Iodine.....	I	126.56	4.95	114	214	Courtois.	1811
31	Iridium.....	Ir	192.65	22.7	1,950		Tennant.	1803
32	Iron.....	Fe	55.91	7.8439	1,800		Known to ancients.	
33	Lanthanum....	La	138.53	6.1	1,000?		Mosander.	1839
34	Lead.....	Pb	206.47	11.4	334	1,600	Known to ancients.	
35	Lithium.....	Li	7.01	0.5385	180		Arvedson; Davy.	1817
36	Magnesium....	Mg	23.96	1.743	750		Davy.	1807
37	Manganese....	Mn	53.91	7.14	1,900		Scheele.	1774
38	Mercury.....	Hg	199.71	13.59	—40	350	Known to ancients.	
39	Molybdenum..	Mo	95.63	8.64	Very high.		Hjelm (Scheele, acid, 1778).	1782
40	Nickel.....	Ni	57.93	8.8	1,600		Cronstedt; Bergmann.	1731
41	Niobium.....	Nb	94.0	7.6	?		H. Rose.	1846
42	Nitrogen.....	N	14.2	14.2		—203	Rutherford.	1772
43	Norwegium....	Ng	214.0	9.441	350		Tellef Dable.	1879
44	Osmium.....	Os	198.49	21.04	2,500		Tennant.	1803
45	Oxygen.....	O	16.9633	15.9633		—203	Priestley; Scheele.	1774
46	Palladium....	Pd	105.74	11.8	1,500		Wollaston.	1803
47	Phosphorus?..	P	32.0?	?	?		Delafontaine	1877
48	Phosphorus†..	P	30.96	1.826	44	290	Brand.	1669
49	Platinum.....	Pt	194.42	22.1	1,600		Wood.	1741
50	Potassium....	K	39.02	0.865	62.5	720	Davy.	1807
51	Rhodium.....	Rh	104.06	12.10	2,000		Wollaston.	1804
52	Rhubidium....	Rb	85.25	1.52	38		Kirchhoff and Bunsen.	1861
53	Ruthenium....	Ru	104.22	11.4	1,800		Claus.	1843
54	Samarium.....	Sm	150.0?		?		L. de Boisbaudran.	1877
55	Scandium.....	Sc	43.96	↑	?		Nilsen.	1879
56	Selenium.....	Se	78.997	4.28	217	685	Berzelius.	1817
57	Silicon.....	Si	28.2	2.49	Very high.		Berzelius.	1823
58	Silver.....	Ag	107.68	10.47	957		Known to ancients.	
59	Sodium.....	Na	23.0	0.972	95.6	960	Davy.	1807
60	Strontium....	Sr	87.37	2.542	490?		Davy.	1807
61	Sulphur.....	S	31.984	2.45	115	447	Known to ancients.	
62	Tantalum.....	Ta	182.14	7.02	?		Hatchet, Ekeberg, 1803	1801
63	Tellurium....	Te	127.96	6.183	455	1,300	Müller v. Reichenstein.	1782
64	Terbium?....	Tb	148.8	?	?		Delafontaine.	1877
65	Thallium.....	Tl	203.72	11.8	290		Crookes; Lamy.	1861
66	Thorium.....	Th	233.41	7.66	?		Berzelius.	1828
67	Thulium?....	Tu	170.7	?	?		Cleve.	1879
68	Tin.....	Sn	117.7	7.29	230	1,600	Known to ancients.	
69	Titanium.....	Ti	49.85	5.3			Gregor.	1791
70	Tungsten.....	W	183.61	19.13	Very high.		Scheele; De Luyart.	1784
71	Uranium.....	Ur	187.25	20.25	?		A. Guyard.	1869
72	Uranium.....	U	238.48	18.33	1,700?		Klaproth(1789); Pellgrot	1841
73	Vanadium....	Vd	51.26	5.5	?		Sefstroem.	1830
74	Ytterbium....	Yb	172.76	?			Marignac.	1878
75	Yttrium.....	Yt	89.82				Gadolín; Ekeberg.	1797
76	Zinc.....	Zn	64.91	6.8	417	927	Known since	1200
77	Zirconium....	Zr	89.37	4.15	Very high.		Klaproth; Troost.	1789

* Discovered hydrofluoric acid.

† Red phosphorus melts at 225° C.; specific gravity, 2.10.

‡ Specific gravity of oxide is 8.86.

§ Specific gravity of oxide is 9.17.

Table of Atomic Weights, $H=1$. Periodic System of Elements.*

I. MONADS.	II. DYADS.	III. TRIADS.	IV. TETRADES.	V. TRIADS AND PENTADS.	VI. DYADS AND HEXADS.	VII. MONADS AND HEPTADS.	VIII.
Li = 7.007	Be = 9.005	B = 10.041	C = 11.074	N = 14.021	O = 15.9633	F = 18.984	$\left\{ \begin{array}{l} \text{Fe} = 55.913 \\ \text{Ni} = 57.928 \\ \text{Co} = 58.987 \\ \text{Cu} = 63.173 \end{array} \right.$
Na = 22.998	Mg = 23.959	Al = 27.003	Si = 28.195	P = 30.958	S = 31.964	Cl = 35.370	
K = 39.019	Ca = 39.990	Sc = 43.980	Ti = 48.000	Vd = 51.256	Cr = 52.009	Mn = 53.906	
(Cu = 63.173)	Zn = 64.905	Ga = 68.854		As = 74.918	Se = 78.797	Br = 79.768	
Rb = 85.251	Sr = 87.374	Yt = 89.816	Zr = 89.367	Nb = 94.000	Mo = 95.888		$\left\{ \begin{array}{l} \text{Ru} = 104.217? \\ \text{Rh} = 104.055 \\ \text{Pd} = 105.737 \\ \text{Ag} = 107.675 \end{array} \right.$
(Ag = 107.675)	Cd = 111.770	In = 113.398	Sn = 117.698	Sb = 119.955	Te = 125.000	I = 126.557	$\left\{ \begin{array}{l} \text{Os} = 188.494? \\ \text{Ir} = 192.651 \\ \text{Pt} = 194.415 \\ \text{Au} = 196.155 \end{array} \right.$
Cs = 132.583	Ba = 136.763	La = 138.526	Ce = 140.424	Di = 146.180	Tb = 148.800(?)		
.....				Er = 165.891			
.....		Yb = 172.760		Ta = 182.144	W = 183.610		
(Au = 196.155)	Hg = 199.712	Tl = 203.715	Pb = 206.471	Bi = 207.523	Ng = 214.000		
.....			Th = 233.414		Ur = 238.482		

* From Curtman's Chemical Analysis.

Symbols and formulæ.

The use of symbols and formulæ forms a chemical short-hand writing. Elements are represented by symbols chosen from initials of Latin names (H, O, N, S, P, C, K, B, I, etc.), or, when more than one begins with the same letter, another characteristic letter is added. The first letter is always a capital, the second is small (lower case): Ag, Al, As, Au; B, Ba, Be, Bi, Br; C, Ca, Cd, Ce, Cl, Co, Cr, Cs, Cu; (Da), Di; E (or Er); F (or Fl), Fe; Ga; H, Hg, Ho; I, In, Ir; K; La, Li; Mg, Mn, Mo; N, Na, Nb, Ng, Ni; O, Os; P, Pb, Pd, Pp, Pt; Rb, Rh, Ru; S, Sb, Sc, Se, Si, Sm, Sn, Sr; Ta, Tb, Te, Ti, Th, Tl, Tu; Ur (Ul); V (or Vd); W; Y (or Yt), Yb; Zn, Zr. Each symbol stands for *one atom* of the element it represents unless a figure is attached to the lower (or upper) right-hand corner, which indicates a higher number of atoms: H = 1 atom of hydrogen, H₂ = 2 atoms of hydrogen.

Formulæ are used to represent the composition of molecules and all of the symbols are written in succession: HCl, a molecule of hydrochloric acid; NH₃, of ammonia; H₂O, of water; BiCl₃, of chloride of bismuth; HNO₃, of nitric acid; H₂SO₄, of sulphuric acid. To multiply the formula, either a large number is written before it: 2H₂O = 2 molecules of water; or it is inclosed in parentheses and the small number follows it: (H₂O)₂ = 2 molecules of water. To represent chemical processes, algebraic signs are used to form equations: H₂SO₄ (sulphuric acid) + Zn = H₂ + ZnSO₄ (sulphate of zinc). In expressing more complicated processes and formulæ, the brackets, [], are used: [(NH₄)₂S]₂.

Frequency of occurrence of elements: In air, O and N; in water, O and H; in earth's crust (estimated):

	Per cent.
O.....	44·00 to 48·70
Si.....	23·80 to 36·20
Al.....	9·90 to 6·10
Fe.....	9·90 to 2·40
Ca.....	6·60 to 0·90
Mg.....	2·70 to 0·10
Na.....	2·40 to 2·50
K.....	0·70 to 3·10
	100·00 100·00

and all other elements in small percentages.

Division of elements.

The elements are divided into *metals* and *metalloids*; these are not strictly separated by different properties; some elements stand on the border between both of the divisions. They are maintained only for convenience of study. Some writers count 25 metalloids (including Sn, Zl, Bi, etc.), others only 13 or 15. Others do not make this division, but adopt numerous smaller ones.

Explanation of *atomicity* or *combining power* (*valency*) as compared with hydrogen.

Classification of elements by atomicities: Monads, H, Cl, I, K, Na, etc.; dyads, O, S, Ca, Ba, Zn, etc.; triads, N, P, Au, Tl, B, etc.; tetrads, C, Si, Ti, Pt, Sn, etc.; pentads, N, Bi, Sn, As, etc.; hexads, Ur, W, Mo, etc.

Variability of atomicity: CO, CO₂; FeO, Fe₂O₃.

Method of marking atomicity: Hⁱ, monad; Oⁱⁱ, diad; Nⁱⁱⁱ, triad; C^{iv}, tetrad; N^v, pentad, etc.

Subdivisions of Metalloids.

Typical Elements.

Hydrogen, H ⁱ	= 1·00	Nitrogen, N ⁱⁱⁱ	= 14·02
Oxygen, O ⁱⁱ	= 15·96	Carbon, C ^{iv}	= 11·97

Chlorine Group.

Fluorine, F ⁱ	= 18·98	Bromine, Br ⁱ	= 79·75
Chlorine, Cl ⁱ	= 35·37	Iodine, I ⁱ	= 126·56

Sulphur Group.

Oxygen, O ⁱⁱ	= 15·96	Selenium, Se ⁱⁱ	= 78·797
Sulphur, S ⁱⁱ	= 31·98	Tellurium, Te ⁱⁱ	= 127·960

Nitrogen Group.

(Nitrogen, N ⁱⁱⁱ)	= 14)	(Antimony, Sb ⁱⁱⁱ)	= 119·96)
Phosphorus, P ⁱⁱⁱ	= 30·96	(Bismuth, Bi ⁱⁱⁱ)	= 207·52)
Arsenic, As ⁱⁱⁱ	= 74·92		

Boron, B ⁱⁱⁱ	= 10·94
Silicon, Si ^{iv}	= 28·20

(Carbon group, C, Si, Ti, Nr, Sn, Th.)

II. HYDROGEN.

Symbol, H. Monad (Hⁱ). Atomic weight, 1. Specific gravity, 1. Boils at about —215° C.

Discovered by Paracelsus, 1500. Described by Cavendish, 1760. Occurs seldom in the free state; mostly in water.

Description: A colorless, inodorous, tasteless gas condensed to a liquid by Pictet and Cailletet, 1877. The lightest gas known. One litre (61·024 cubic inches, 2·1135 wine, or 1·76 imperial pints) at 760 millimetres pressure and 0° C. weighs 0·089586 grammes. One cubic metre = 89·685 grammes. One gallon = 4·94 grains. Air being taken = 1, the specific gravity of H is 0·0693. H is 14·442 times lighter than air; 11,167 than water; 151,803 than mercury.

	Grammes.
1 litre of hydrogen.....	= 0·089586
1 " air.....	= 1·293000
1 " water.....	= 1,000·000000
1 " mercury.....	= 13,598·000000

A room 48x34x12 feet = 15,504 cubic feet = 439 cubic metres; filled with H would contain 39·328 kilos (86·522 avoirdupois pounds); with air, 567·975 kilos (1,249·545 avoirdupois pounds). H is used for filling balloons. In transferring it from one vessel to another it will rise

so as to be held in glass vessel mouth down. H refracts light $6\frac{1}{2}$ (6,614 : 1,000) times stronger than air. In air it is inflammable, burning with a pale, non-luminous hot flame, forming water.

Experiments with hydrogen: Collodion balloons; igniting H in closed bell-glass, candle does not burn in it, gas burns at bottom; igniting at open jet of bell-glass, light gas rises, burns, and residuary mixture with air explodes; chemical harmonica; mixture with oxygen explosive (gas pistol); high temperature of combination; compound blow-pipe; Drummond light (Holyhead, 1826).

Flame Temperatures.

	Centigrade.	Fahrenheit.
Hydrogen burning in oxygen.....	4460	8061
Sulphur " "	993	1820
Marsh gas " "	1288	2350
Carbon disulphide burning in oxygen	1202	2195
" monoxide "	1673	3042
Hydrogen in air.....	1788	3259

Preparation of hydrogen: Decomposition of water by K or Na, $2\text{H}_2\text{O} + \text{K}_2 = 2(\text{KOH}) + \text{H}_2$; decomposition of water by red hot iron, $2\text{Fe} + 3\text{H}_2\text{O} = \text{Fe}_2\text{O}_3 + 3\text{H}_2$; decomposition of sulphuric or hydrochloric acid by Zn; $\text{H}_2\text{SO}_4 + \text{Zn} = \text{ZnSO}_4 + \text{H}_2$; $2\text{HCl} + \text{Zn} = \text{ZnCl}_2 + 2\text{H}$; decomposition of water by electrolysis, explanation of apparatus [Fig. 101], displacement of water by gas in receiver.

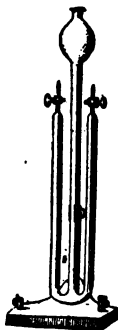


FIG. 101.

One and one-half per cent by volume is soluble in water.

Occlusion of H by Pd: Graham's experiments; iron; meteorite of Lenarto contains occluded H; Graham's *hydrogenium* metal predicted from properties of Pd alloy; 800 to 900 per cent of H occluded in Pd, which increases 1.6 per cent; density of solid H predicted as 0.715 (first, 1.951).

Spectrum of hydrogen: C. F. near G. (h); stellar spectra, solar atmosphere, and nebulae contain H.

H is dia-magnetic and electro-positive.

Solid and liquid hydrogen: Cailletet and Pictet, 1877, condensed H to a liquid, boiling at about -203° C. (and at -140° C. under 600 atmospheres pressure). It is a colorless, transparent liquid, solidifies by its own evaporation.

H is the vapor of a metal. Critical temperature. Absolute boiling point. Action of Pt sponge on H (condenser of O). Dæbereiner's lamp [Fig. 102]; union thus occurring at lower temperature than of O and H ordinarily.



FIG. 102.

Diffusion of gases: Hydrogen, having least specific gravity, is most diffusible; gases are diffusible in an inverse ratio of square root of densities, $H:O = \sqrt{16}:\sqrt{1}$, or $H = 4$ and $O = 1$; caution in keeping H in closed apparatus, as it mixes by diffusion and makes explosive mixtures of O and H.

Compounds of hydrogen: Called *hydrides*; not as many as O forms; ClH , OH_2 , NH_3 , CH_4 , PH_3 , AsH_3 , SbH_3 , Cu_2H_2 , SiH_4 , SH_2 , Pd_2H , etc.

Reduction of metals from their oxides by hydrogen: Fe, Cu, etc., by uniting with O to form H_2O (water); with electro-negative elements, or groups of elements, it forms acids.

Hydrogen forms acids whose H is replaceable by metals: HCl , HI , HBr , HF , H_2S , H_2SO_4 , H_2CO_3 , H_3PO_4 , H_3BO_3 , etc.

Monobasic, dibasic, and tribasic acids. Monatomic element.

III. OXYGEN.

Symbol, O. Diad (O^{II}). Atomic weight, 15.9633. Specific gravity, 15.9633 ($H = 1$). Boils at about -203° C.

Discovered in England by Priestley, 1774 (centennial celebration at his grave, Northumberland, Pa., August 1, 1874), and in Kœpfer by

Scheele. Named oxygen by Lavoisier, 1778; dephlogisticated air by Priestley; empyreal air by Scheele (Stahl and Becker's phlogiston).

Occurs in air ($\frac{1}{5}$); water ($\frac{8}{9}$); solid crust of earth (nearly $\frac{1}{2}$).

	Air by weight.	Air by volume.
N.....	76.8	78.492
O.....	23.2	20.527
H ₂ O.....		0.840
CO ₂		0.041
	100.0	100.000

Physical properties: Colorless, tasteless, odorless gas; condensed to liquid by Cailletet, 1877, and by Pictet (500 atmospheres pressure at $-1,400^{\circ}$ C.); less refractive than air, index = 0.86161; 1 litre weighs 1.429802 grammes at 760 millimetres pressure and 0° C., 1 gallon = 79.09 grains, 1 cubic foot = 592 grains, 12 cubic feet = 1 pound; air = 1, oxygen = 1.10568.

Chemical properties of oxygen: Thirty-seven parts soluble in 1,000 parts of water by volume (1,000 volumes of water dissolves 15 volumes of N); 10 volumes of O dissolves in 1 of melted silver, and is expelled on solidification; O condenses on platinum sponge and black; does not burn in air, but occasions the combustion of other elements (H, C, S, P, etc.) by uniting with them to form oxides; only fluorine does not combine with O (?) (oxyfluoride of copper, CuOF, oxyfluoride of cuprammonium and fluo-benzoic acid); the chemical union is initiated at various temperatures and produces heat.

Experiments with oxygen: Rekindling glowing chip; burning charcoal, camphor, S, P, Fe, steel, K, and Mg.

Increase of weight of body by burning with O. Different reactions with test paper. Some acid, some alkaline oxides. Oxides deprived of O are *reduced*. Definite quantities of elements uniting with O produce definite amounts of heat.

1 pound of charcoal.....	heats	78.0	pounds water from	0° C. to 100° C.
1 " bituminous coal.....	"	60.0	" " "	0° C. to 100° C.
1 " alcohol.....	"	67.5	" " "	0° C. to 100° C.
1 " olive oil.....	"	90.0	" " "	0° C. to 100° C.
1 " hydrogen.....	"	236.0	" " "	0° C. to 100° C.

One pound of O, with a sufficient quantity of H to consume it, heats:

29.5 pounds water from.....	0° C. to 100° C.
With charcoal, 29.0 pounds water from.....	0° C. to 100° C.
" alcohol, 28.0 " " " " " " " " " "	0° C. to 100° C.
" ether, 28.5 " " " " " " " " " "	0° C. to 100° C.

Consumption of O by animals and reproduction by vegetables. Action of chlorophyl and actinic rays. An adult consumes 46,087 cubic inches (15,661 grains, over 2 pounds) of O in twenty-four hours, forming 14,930 cubic inches, or 8,584 grains ($1\frac{1}{4}$ pounds), carbon dioxide. This amounts to 746 pounds in one year (837 pounds, Menzler).

Animal heat. Heat of slow oxidation, cremacausis.

To initiate oxidation, various elements require different temperatures: P, alkalies and alkali earths unite with O at ordinary temperature of air; dry Fe requires a high heat; Zn, about $260^{\circ}\text{C}.$; charcoal, a red heat ($500^{\circ}\text{C}.$); H, about $320^{\circ}\text{C}.$ Moisture and minute subdivision favors oxidation of some substances. Different degrees of oxidation of same element: Hg_2O , black mercurous oxide; HgO , red mercuric oxide; H_2O , H_2O_2 ; MnO MnO_2 . Different names given to distinguish oxides:

<i>Mon(o)</i> oxide.	<i>Sesquioxide.</i>
<i>Dioxide.</i>	<i>Suboxide.</i>
<i>Trioxide.</i>	<i>Hyper(or per)oxide.</i>
<i>Tetr(a)</i> oxide.	— <i>ic</i> oxide.
<i>Pent(a)</i> oxide.	— <i>ous</i> oxide.

The above nomenclature is also applied to chlorides, etc.

Nomenclature of acids and salts:

—*ic* acid produces —*ate* salt.

—*ous* acid produces —*ite* salt.

Hyper or *per*—*ic* acid produces *Hyper*—*ate* salt.

Hypo—*ic* acid produces *Hypo*—*ate* salt.

Hypo—*ous* acid produces *Hypo*—*ite* salt.

Example: Nitrogen oxides and acids.

Preparation of oxygen: 2HgO (red oxide of mercury) + heat decomposes into $\text{Hg}_2 + \text{O}_2$ (molecules of free $\text{O} = \text{O}_2$); 3MnO_2 + heat = $\text{Mn}_3\text{O}_4 + \text{O}_2$; 2KClO_3 + heat = $2\text{KCl} + 3\text{O}_2$, yields O at lower temperature by admixture of MnO_2 ; 2 stages of reaction; 2KClO_3 + heat = $\text{KClO}_4 + \text{KCl} + \text{O}_2$, then $\text{KClO}_4 = \text{KCl} + 2\text{O}_2$; MnO_2 or $\text{K}_2\text{Cr}_2\text{O}_7 + \text{H}_2\text{SO}_4$ (sulphuric acid); first made from saltpeter (Gehler, 1798); electrolysis of water: 2 volumes of H + 1 volume of $\text{O} = \text{H}_2\text{O}$; decomposition of CO_2 by green leaves and light; numerous other processes, KNO_3 , H_2SO_4 , BaO_2 ; by concentrating O from air in watery solution, etc.

Free O molecule = O_2 . *Status nascendi.* O more active while in separate atoms, before they form into molecules.

Diatomic element ($\text{O} = \text{O}$); requires 2H^1 , or 2 atoms of monatomic K^1 , or Na^1 , or Ag^1 ; unites with 1 atom of Hg^{11} ; $\text{H} - \text{O} - \text{H}$, $\text{K} - \text{O} - \text{K}$, $\text{K} - \text{O} - \text{H}$, $\text{Ag} - \text{O} - \text{Ag}$, $\text{Ca} = \text{O}$, $\text{Zn} = \text{O}$, $\text{O} = \text{C} = \text{O}$.

IV. OZONE.

Formula, O_3 . Molecular weight, 48. Specific gravity, 24 ($\text{H} = 1$).

Source: When O is set free at low temperatures, or when silent electric discharges are sent through common O, or when P is oxidized at low temperatures, etc., a modification (allotropic) of common O is formed, having 3 atoms instead of 2 in a molecule. Schoenbein, of

Basel (1840), discovered it by electrolysis of water. Name from $\delta\zeta\omega$, I smell.

Properties of ozone: It is easily decomposed, and thus becomes nascent O of great affinity. Heat at 100° C. slowly, at 290° C. instantly, decomposes it. It is produced by green leaves, hence a frequent constituent of the air. It liberates I from iodide of potassium (KI) as shown by starch. *Ozometry.* O_3 is irritating to the organs of smell and respiration. It oxidizes Pb, Ag, etc., at ordinary temperatures. Found in ether and ethereal oils. It colors strychnine and guaiacum. Sparingly soluble in water. Not obtained pure, but only reduced (5 to 6 per cent) with inactive O.

Ordinary oxygen molecule is $O = O$. Ozone molecule is .

Blue color of gas shown in tube 1 metre long. Indigo color when compressed. Easier to condense to a liquid than ordinary O. Rapidly absorbed by oxalic acid in aqueous solutions.

Ozonides: Permanganic acid, chromic acid, lead dioxide, etc. Obtained by adding sulphuric acid to potassium manganate or permanganate.

V. WATER.

Formula, H_2O . Molecular weight, 18.

Natural water is never entirely pure; salts dissolved in spring, well, or river water. Rain and snow water contains gases, etc. Purified by distillation. Formed by union of 2 volumes H and 1 volume O, the mixture burns with explosion when ignited by electric spark or flame. Union of the two gases also occurs by finely divided Pt, which becomes incandescent in the mixture. Two grammes H burning with 15.9633 grammes O develops 67.4 calories. The flame may be made to exceed $2,000^{\circ}$ C. when H is passed over cupric or other metallic oxides. The H unites with the O to form H_2O and reduces the oxide to a metal. H. Cavendish (1781) by synthesis first ascertained composition of H_2O .

Properties of water: Colorless, tasteless, inodorous liquid between 0° C. and 100° C. Greatest density at 4° C. Freezes at 0° C., unless cooled under compression, when it remains liquid to -24° C. Boiling point depends on pressure, etc.

Expansion of H_2O on freezing: 100 cubic centimetres $H_2O = 1$ gramme at 4° C. = 0.9 gramme at 0° C. Latent heat of liquid $H_2O = 79$ calories; of steam = 536 calories. Tension vapor of H_2O at

Millimetres.		Millimetres.	
-20° C.	= 00.927	60° C.	= 148.791
0° C.	= 04.600	80° C.	= 354.280
20° C.	= 17.391	90° C.	= 525.450
40° C.	= 54.901	100° C.	= 760.000

or 1 atmosphere; above boiling point:

	Atmospheres.		Atmospheres.
111.7° C.....	= 1.5	159.2° C.....	= 6
120.6° C.....	= 2.0	170.8° C.....	= 8
133.9° C.....	= 3.0	180.3° C.....	= 10
144.0° C.....	= 4.0	213.0° C.....	= 20

One volume of H_2O expands to 1,696 volumes of steam at 100° C.; 1 litre of steam weighs 0.8064 gramme. Condensing H_2O vapor forms hollow vesicles filled with air. Clouds, fogs, etc. Great solvent powers of H_2O , usually increased by heat. Crystals formed by cooling a hot saturated solution (lime, Di, etc., exceptions).

Hygroscopic substances; efflorescence; water of crystallization: $Na_2CO_3 + 10H_2O$ has 63 per cent of H_2O of crystallization; $MgSO_4 + 7H_2O$ has 51 per cent; and $Al_2(SO_4)_3 \cdot K_2SO_4 + 24H_2O$, 45.5 per cent. Solubility of salts in H_2O at different temperatures:

100° C., 100 parts H_2O dissolves.....	250 parts nitrate of potassium.
0° C., 100 " H_2O ".....	10 " "
100° C., 100 " H_2O ".....	39 " chloride of sodium.
0° C., 100 " H_2O ".....	36 " "
100° C., 460 " H_2O ".....	1 part sulphate of calcium.
0° C., 400 " H_2O ".....	1 " "

Natural waters.

Rain or snow contains only gases, ammonium, nitrous and nitric acids as impurities, unless precipitated through dusty atmosphere.

River water contains suspended and dissolved impurities, as clay, silt, etc. The salts in solution are $CaH_2(CO_3)_2$, Mg, Fe, NaCl, $CaSO_4$, etc. Sewage and other contaminations, micro-organisms, ova of entozoa, disease germs, moulds, bacteria, etc. *Well and spring water* generally contains more carbon dioxide; hence, in limestone regions more Ca, etc., causing hard water. Temporary hardness caused by acid carbonate of calcium; permanent hardness by sulphate of calcium; means of softening hard water; soap test for hard water (10 grammes dry soap dissolved in 500 grammes alcohol, then diluted with 1,000 cubic centimetres water). All of the waters are generally included among *potable waters*.

Analysis of waters: One gallon contains of dry residue:

	Grains.		Grains.
Thames.....	18 to 20	Nile (according to season).....	9 to 22
Rhine, at Basle.....	12	Seine, at Paris.....	20
Spree, at Berlin.....	8		

Suspended and dissolved matter in Mississippi water at St. Louis:

	Grains.		Grains.
Summer.....	171	Winter.....	94

Organic residue per gallon:

	Grains.		Grains.
Summer.....	2½ to 3	Winter.....	4 to 5

Total residue per gallon filtered:

	Grains.		Grains.
Summer.....	33	Winter.....	13

Mineral waters is the name applied when constituents not ordinarily found in well water are present, or when large residues remain. *Thermal waters* come from great depths. They contain Si and various constituents: Hot Springs (Ark.); Geysers of Iceland, Yellowstone Park, Auckland, etc. *Carbonated waters* contain carbon dioxide in great abundance, often with alkalies, alkaline earth, Fe, etc.: Selters, pyrmont, vichy, etc. *Saline waters* contain sulphates or chlorides of Na, K, Li, Mg, Ca, etc. *Sulphurous waters* contain sulphureted hydrogen or alkaline sulphides or oxysulphides: Kiebourne, Harrowgate, White Sulphur Springs (Va.), Searcy (Ark.), Belcher (St. Louis), etc. *Chalybeate waters* contain an abundance of Fe. Popular prejudices in regard to mineral and other drinking waters. Analysis; difficulties; evaporation and ignition; permanganate process; ammonia process of Wanklin; combustion process of Franklin; determination of single constituents is very laborious. *Sea water* from British Channel, specific gravity 1.0274, contains in 1,000 parts (Schweitzer):

Water	964.745	Sulphate of magnesium.....	2.206
Chloride of sodium.....	27.059	“ of calcium.....	1.406
“ of potassium.....	0.788	Carbonate of calcium.....	0.033
“ of magnesium.....	3.666	Iodine.....	traces
Bromide of	0.029	Ammonia.....	“

Sea water freezes at -21° C. Use in refrigerators in connection with ice machine.

Purification of water: Filtering to remove suspended matter; sand, charcoal (which removes dissolved matter), carbon, silica filters, etc.; adding *potassium permanganate* to destroy organic matter; *lime-water* to decompose calcium bicarbonate; *alum* for precipitating suspended matter; *distilling*, precautions.

Influence of water on climate; circulation of water on the earth; water in organic tissues: Necessity to vital operations; *plants* take up water by the roots and exhale by leaves; *animals* take it into the stomach and exhale it by the lungs, skin, kidneys, etc.

Water in food (solids):

Substance.	Per cent of H ₂ O.	Substance.	Per cent of H ₂ O.
Albumen.....	88	Beefsteak (cooked).....	69
Yolk of egg.....	50	Chicken.....	79
Oysters.....	87	Potatoes.....	75
Beefsteak (raw)...	75	Cucumbers (green).....	96

(liquids):

Beer (extract, 8; alcohol, 8).....	86	Cellulose (linen).....	40
Milk.....	92	“ (wood).....	44

Osmosis: In vessel containing solution, *exosmotic* current from solution to pure H_2O is represented by “—;” *endosmotic* current from pure H_2O to solution by “+.”

Oxalic acid	— 148	Chloride of sodium	+ 2
Hydrochloric acid	— 92	Nitrate	+ 2
Chloride of gold	— 54	Chloride of potassium	+ 18
Mercuric chloride	— 3	Sulphate	+ 21
		Mercuric chloride	+ 121
		Acid phosphate of sodium	+ 311
		Chloride of aluminium	+ 540

Baths in pure H_2O , and in saline solutions; influence on tissues, loss and gain of salts, etc.

V. HYDROGEN DIOXIDE.

Formula, H_2O_2 ($H—O—O—H$). Molecular weight, 34. Specific gravity, 1.457.

H_2O_2 is formed in small quantity when O_3 is formed in the presence of H_2O (autozone).

Preparation of H_2O_2 : Barium dioxide decomposed by dilute acids (1 part sulphuric acid + 5 parts H_2O).

Properties of H_2O_2 : When concentrated, its aqueous solution is colorless, inodorous, has a bitter taste, does not freeze at $-30^\circ C.$, of syrupy consistency; easily decomposed by heat when concentrated by Pt powder, Au, or Ag, by some metallic oxides (HgO , Ag_2O , MnO_2 , etc.), $H_2O_2 + Ag_2O = H_2O + O_2 + Ag_2$; potassium permanganate solution (with sulphuric acid) is dissolved and free O , H_2O , and manganous sulphate forms; iodide of potassium is not affected by H_2O_2 (difference from O_3) unless ferrous sulphate is present; chromic acid is converted into blue perchloric acid, soluble in ether.

Hydroxyl, $—O—H$, monatomic, unsaturated group.

VI. NITROGEN.

Symbol, N. Triad and pentad (N^{III} and N^V). Atomic weight, 14.021. Specific gravity, 14.021 (air = 1, specific gravity is 0.969).

Occurs in the free state mingled with O in the air: N 76.8 by weight, 78.492 by volume; O 23.2 by weight, 20.627 by volume.

Described by Rutherford, 1772. Called *azotum* by Lavoisier, and *nitrogenium* by Chaptal. Occurs in compounds (niter, ammonium salts), animal and vegetable tissues, meteoric iron, etc.

Preparation: From air by removing O and carbonic acid gas by combustion, etc.; from ammonia by removing H by Cl; from ammonium nitrite by heating (17 parts potassium nitrate + 13 parts ammo-

niium sulphate dissolved in water + concentrated aqueous solution of potassium dichromate to color red, heated evolves N).

Description: N is a colorless, inodorous, tasteless gas, condensed by pressure and cold to a liquid (Pictet and Cailletet, 1877); soluble in H_2O (1.5 per cent by volume); not combustible; burning bodies are extinguished in it; irrespirable only from want of O; slightly lighter than air, 1 litre weighs 1.2544 grammes; unites with H, with O, and some other elements (B, Si, Ti, Mg, C) forming *nitrides*; is mostly *triatomic*, but sometimes *pentatomic*, etc.

Air is a mixture of gases: 100 cubic inches = 30.82926 grains, 1 cubic foot = 1.2 avoirdupois ounces; a room 14x20x50 feet = 1,000 avoirdupois pounds.

	Specific gravity.		Grammes.
Hydrogen	= 1.000, 1	cubic metre =	89.68
Air	= 14.442, 1	" "	= 1,293.00
Water	= 11,167.000, 1	" "	= 1,000,000.00
Mercury	= 151,803.000, 1	" "	= 13,598,000.00

Composition of air determined by *eudiometer*, either by combination of O with H, absorption by pyrogallate of potassium, by chloride of copper, or by P.

VII. AMMONIA.

Formula, NH_3 . Molecular weight, 17. Specific gravity, 8.5 (air, 1, specific gravity = 0.5893).

Formed by silent electric discharges through a mixture of 1 volume N and 3 volumes H; more readily in presence of H_2O , or sulphuric acid. During thunder storms ammonium nitrite and nitrate form. It is also formed by reduction of nitric acid by *nascent* H, as when treating nitrates with Al, Zn, Fe, caustic alkalies; during decomposition by putrefaction of nitrogenous animal and vegetable tissues, or by heating them with caustic potassa or even alone; hence the name *hartshorn* spirits. The name ammonia from place of preparation.

Preparation of NH_3 : Ammoniacal salts are derived from gas liquors and these are distilled with quick-lime.

Properties of NH_3 : Colorless gas; peculiar, pungent odor; soluble in H_2O in different proportions according to temperature; at $0^\circ C$. 1 cubic centimetre H_2O dissolves 0.899 gramme NH_3 , the saturated solution has specific gravity = 0.875 and contains about 1,149 volumes; at $20^\circ C$. 1 cubic centimeter H_2O dissolves 0.520 gramme = 681 cubic centimetres NH_3 ; at $100^\circ C$. all NH_3 is expelled from aqueous solutions; NH_3 contains 82.39 per cent N and 17.61 per cent H; gas and solution shows alkaline reaction; when brought near hydrochloric acid solution white fumes of ammonium chloride are formed; does not burn in air, but in pure O it may be ignited and gives a yellow flame, feebly luminous, forming N and H_2O ; when heated or electro-

lyzed 2 volumes are decomposed into 1 volume N and 3 volumes H; it is a strong base, neutralizing acids, forming salts.

Saturated solutions vary in strength according to the temperature: at 10° C. specific gravity = 0.875; at 0° C. specific gravity = 0.870 and contains 46 to 47 per cent NH_3 .

Experiments with NH_3 : Solubility in H_2O shown by filling flask with gas, into which H_2O rises as a fountain jet; alkaline reaction shown by adding phenolphthaleine, colored red by NH_3 .

Aqua ammoniæ (U. S. P.) contains 10 per cent by weight of NH_3 gas; specific gravity, 0.959 at 15° C.; 8.5 grammes = 8.9 cubic centimetres, and neutralizes 50 cubic centimetres normal acid solution. By Baume's hydrometer it marks 16°.

Aqua ammoniæ fortior (U. S. P.) contains 28 per cent by weight of NH_3 gas; specific gravity, 0.900 at 15° C. = 26° B.

Ammonium (NH_4), *amidogen* (NH_2), *imidogen* (NH), and *hydroxylamine* (NH_2OH), to be treated of hereafter.

Liquid anhydrous NH_3 : NH_3 gas condenses at ordinary temperature by pressure of 6 to 7 atmospheres, or at -40° C. without pressure; it is a colorless liquid, specific gravity = 0.6234 at 0° C.; crystallizes at -75° C. to a white, transparent solid (solidifies by its own evaporation); when brought in contact with H_2O it hisses like red-hot Fe; used in artificial production of ice; condensing pressures of NH_3 at:

	Atmospheres pressure.
0° C.....	= 4.4
20° C.....	= 8.0
40° C.....	= 15.5

Experiments with liquid NH_3 : Glass model of Carré's NH_3 ice-machine; dropping liquid NH_3 into H_2O ; Na dissolves in anhydrous NH_3 with blue color.

Tests for NH_3 : Alkaline reaction of vapors; deep blue color with cuprous sulphate; white fumes with hydrochloric and other volatile acids; Nessler's test, brown color of iodide of tetra-mercur-ammonium, mercuric-potassium iodide; Curtman's test, brown color with solution of chloral hydrate in sulphureted hydrogen water.

Oxides of nitrogen:

1. N_2O monoxide.....	28:1 × 16
2. NO dioxide (N_2O_2).....	28:2 × 16
3. N_2O_3 trioxide.....	28:3 × 16
4. NO_2 tetroxide (N_2O_4).....	28:4 × 16
5. N_2O_5 pentoxide.....	28:5 × 16

VIII. NITROGEN MONOXIDE.

Formula, N_2O . Molecular weight, 44. Specific gravity, 22 (air = 1, specific gravity is 1.527). Boils at -88° C. Synonyms: Nitrous oxide, laughing gas (nitrosyl).

Properties: Similar to O; glowing chip rekindled, P burns, etc.; sweetish odor and taste; soluble in H₂O (1 volume H₂O dissolves 0.608 volumes of N₂O at 21° C., or 1.305 volumes at 0° C.); colorless gas; contains 63.77 per cent N and 36.23 per cent O; coercible at 0° C. by 50 atmospheres pressure, and at -88° C. by normal atmospheric pressure; liquid not miscible with H₂O; mixtures with H explode by electric spark; not capable of sustaining respiration; intoxicating, anæsthetic; heat decomposes it into N₂ and O; passed over hot K it forms K₂O and N₂; does not directly unite with O; passed over hot potassium hydrate (KOH) it forms potassium nitrate (KNO₃) and NH₃; (N₂O + KOH + H₂O = KNO₃ + NH₃); it gives greater heat of combustion than O as there is less difficulty in separating the atoms in



than O=O; with H₂O it forms *nitrosylic* (HNO) or *hyponitrous acid* (N₂O + H₂O = 2HNO); not formed by direct union of H₂O with N₂O, but prepared by action Na amalgam on nitrite or nitrate of potassium; neutralizing by acetic acid and precipitating with nitrate of silver, light-yellow *hyponitrite* of silver (AgNO) is formed, not altered by light but explodes when heated above 110° C., and is soluble in sulphuric acid but reprecipitated by NH₄OH; a concentrated aqueous solution of HNO may be obtained by decomposing AgNO with hydrochloric acid; it can not be isolated without decomposition; is a colorless acid; colors test papers strongly; liberates I from iodide of potassium and reduces potassium permanganate.

Preparation of N₂O: By heating ammonium nitrate (NH₄NO₃); (NH₄NO₃ heated = N₂O + 2H₂O); dissolving Zn in dilute nitric acid (HNO₃); (4Zn + 10HNO₃ = 4[Zn(NO₃)₂] + N₂O + 10H₂O); it is received over water of 30° C.

Experiments with N₂O: Rekindling glowing chip; burning P; burning S extinguished unless highly heated.

IX. NITROGEN DIOXIDE.

Formula, NO; (N₂O₂). Molecular weight, 29.984. Specific gravity, 14.99 (air = 1, specific gravity is 1.039). Synonyms: Nitric oxide, nitroxyl.

Properties: Colorless gas; contains 46.67 per cent N and 53.33 per cent O; very little soluble in H₂O; condensed to a liquid by pressure of 104 atmospheres at -11° C.; dissolves readily in aqueous solutions of ferrous salts with reddish-brown color, heat again expels the gas; soluble in nitric acid; gives, according to concentration, brown, yellow, green, or blue color, forming N₂O₃; (2HNO₃ + 4NO = 3N₂O₃ + H₂O); is converted into nitric acid by potassium permanga-

nate; $10\text{NO} + 6\text{KMnO}_4 + 9\text{H}_2\text{SO}_4 = 10\text{HNO}_3 + 8\text{K}_2\text{SO}_4 + 6\text{MnSO}_4 + 4\text{H}_2\text{O}$; sustains combustion much less than O or N_2O on account of greater difficulty of separating the O from $\text{N}=\text{O}$; P burns in it, S does not; when heated with H and treated with spongy Pt or heated pumice powder, NO forms NH_3 and H_2O ; $2\text{NO} + 5\text{H}_2 = 2\text{NH}_3 + 2\text{H}_2\text{O}$; mixed with carbon disulphide, it burns with bright blue flame, rich in actinic rays; passed over incandescent Cu, it yields its O; it readily takes up O, and is converted into NO_2 (or N_2O_4), which has reddish-brown color, deepening on heating; hence NO is a reagent for O; it does not change litmus color.

Preparation of NO: By reduction of nitric acid by Cu, sulphurous acid or other reductents (ferrous salts, Hg, etc.), $8\text{HNO}_3 + 3\text{Cu} = 2\text{NO} + 3[\text{Cu}(\text{NO}_3)_2] + 4\text{H}_2\text{O}$ or $4\text{HNO}_3 + 3\text{Ag} = \text{NO} + 3\text{AgNO}_3 + 2\text{H}_2\text{O}$.

The third atomicity of N in $-\text{N}=\text{O}$ being in abeyance makes it combine easily with monads: NOBr, *nitroxyl bromide*; NOCl, *nitroxyl chloride*.

X. NITROGEN TRIOXIDE.

$$\begin{array}{c} \text{N}=\text{O} \\ \diagup \quad \diagdown \\ \text{Formula, } \text{N}_2\text{O}_3; \quad \text{O.} \quad \text{Molecular weight, 75.93. Specific gravity,} \\ \diagdown \quad \diagup \\ \text{N}=\text{O} \end{array}$$

37.96 (air = 1, specific gravity is 2.63). Synonym, nitrous (acid) anhydride. Boils at 2°C .

Properties: Dark-blue liquid; contains 36.84 per cent N and 63.16 per cent O; when boiled it decomposes into NO and NO_2 , which recombine on cooling below -30°C .; decomposed by H_2O into HNO_3 and NO (first HNO_2 ?); $3\text{N}_2\text{O}_3 + 3\text{H}_2\text{O} = 6\text{HNO}_3$, ($= 2\text{HNO}_3 + 4\text{NO} + 2\text{N}_2\text{O} + 2\text{H}_2\text{O}$), or $3\text{N}_2\text{O}_3 + \text{H}_2\text{O} = 2\text{HNO}_3 + 4\text{NO}$.

Preparation of N_2O_3 : Cooling a mixture of 1 volume NO and 1 volume NO_2 to 0°C ., or mixing 4 volumes NO with 1 volume O and cooling; 5 volumes shrink to 2 volumes: 2NO (4 volumes) + O (1 volume) = N_2O_3 (2 volumes); prepared by gently heating nitric acid with arsenious oxide, $2\text{HNO}_3 + 2\text{H}_3\text{AsO}_3 = \text{N}_2\text{O}_3 + 2\text{H}_3\text{AsO}_4$, an orange-colored gas is evolved which condenses to a blue liquid at 0°C .; by heating nitric acid with starch.

XI. NITROUS ACID.

$$\begin{array}{c} \text{N}=\text{O} \\ | \\ \text{Formula, } \text{HNO}_2; \quad \text{OH} \end{array} \quad \text{Monobasic.}$$

It is so unstable as to be unknown in the free state. Forms salts called *nitrites* by exchange of H for a metal: KNO_2 , NaNO_2 , AgNO_2 . It is *monobasic*, having but one exchangeable H atom. May exist in

watery solution when nitric acid is present. Liberates I from iodide of potassium. Colors yellow a solution of *meta-di-amido-benzol*, $[C_6H_4(NH_2)_2]$ in sulphuric acid; this reaction discovers even small traces of HNO_2 (also called phenylene-di-amine). Another test is by adding to nitrites a solution of *sulph-anilic acid* in dilute sulphuric acid, and, after a few minutes, a solution of *naphthyl-amine sulphate*, a red color is produced (HNO_2 is found in the saliva by these delicate tests). Co and K solutions precipitate nitrites (Fisher's salt).

XII. NITROGEN TETROXIDE.

Formula, NO_2 ; (N_2O_4) ; $\begin{array}{c} N=O \\ \diagdown \\ O \end{array}$.

Molecular weight, 46. Specific gravity of gas, 22.9733 (air = 1, specific gravity is 1.59); of liquid, 1.45 ($H_2O = 1$).

Properties: A reddish-brown pungent gas above $22^\circ C.$, but yellow liquid below $22^\circ C.$ which freezes at $-12^\circ C.$, forming colorless crystals; the higher the temperature the more N_2O_4 will dissociate into red NO_2 ; bases convert N_2O into nitrates and nitrites; hot H_2O converts it into nitric acid and NO ; $3NO_2 + H_2O = 2HNO_3 + NO$; use in manufacture of sulphuric acid; lead chamber crystals are $SO_2 \begin{array}{c} \diagup OH \\ \diagdown NO_2 \end{array} + H_2O = H_2SO_4 + HNO_2$, formerly called hyponitric acid; specific gravity of N_2O_4 vapor at $26.7^\circ C. = 2.65$, at $150^\circ C. (NO_2) = 1.58$; NO_2 contains 30.44 per cent N and 69.56 per cent O.

Preparation of NO_2 : Mixing 2 volumes NO with 1 volume O ; reducing nitric acid by light, heat, etc.; heating nitrate of lead, $Pb(NO_3)_2$ heated = $PbO + O + 2NO_2$.

Compounds of NO_2 : A monatomic group; replaces H in organic compounds forming *nitro* derivatives; combines with bases to form nitrates and nitrites, formerly thought to be hyponitrates, $2PbN_2O_5 = 2Pb \begin{array}{c} \diagup NO_2 \\ \diagdown NO_2 \end{array}$ or = $Pb(NO_3)_2 + Pb(NO_2)_2$; chlorine compound is *nitroxyl chloride*, NO_2Cl , a yellow liquid boiling at $-5^\circ C.$

Relation of NO_2 and N_2O_4 : Dissociation at even moderate temperature, $\begin{array}{c} N=O \quad O=N. \\ \diagdown \quad \diagup \\ O-O \end{array}$.

XIII. NITROGEN PENTOXIDE.

Formula, N_2O_5 ; $\begin{array}{c} N \quad N \\ || \quad || \\ OO \quad O \quad OO \end{array}$. Molecular weight, 107.86. Synonym, nitric (acid) anhydride.

Properties: Colorless, rhombic prisms; melts at 30°C. , boils at 47°C. ; decomposes readily (explosive); oxidizes S rapidly; ignites and oxidizes P; unites with H_2O to nitric acid, $\text{N}_2\text{O}_5 + \text{H}_2\text{O} = 2\text{HNO}_3$; unites with nitric acid to unstable $\text{N}_4\text{O}_{11}\text{H}_2 = \text{N}_2\text{O}_5 + 2\text{HNO}_3$, forming a colorless liquid of 1.642 specific gravity at 18°C. , crystallizes at -5°C. ; N_2O_5 contains 25.9 per cent N and 74.1 per cent O.

Preparation: When Cl is passed over silver nitrate at 60°C. , or when pure nitric acid (63 parts) is mixed with phosphorous pentoxide (71 parts) at a low temperature, N_2O_5 is formed.

XIV. NITRIC ACID.

Formula, HNO_3 ; $\begin{array}{c} \text{N} \equiv \text{O} \\ | \\ \text{O} - \text{H} \end{array}$. Monobasic. Molecular weight, 62.91. Specific gravity, 1.54. Boils, 86°C. Synonym, aqua fortis.

Properties: Colorless, pungent liquid; crystallizes at -40°C. ; decomposes readily by heat or light at ordinary temperature into N_2O_4 and O and H_2O , $= 2\text{HNO}_3$; very corrosive; miscible with H_2O in all proportions; when dilute acid is distilled a very watery acid is obtained first until the residue contains 68 per cent HNO_3 of specific gravity 1.414 at 15°C. , boiling at 120.5°C. ; this residue results when strong acid is distilled, very strong passes over first until the 68 per cent remains; this 68 per cent HNO_3 is the commercial strength acid.

Acidum nitricum (U. S. P.) contains 69.4 per cent absolute HNO_3 and 30.6 per cent H_2O , specific gravity 1.42, $= \text{HNO}_3 + \text{H}_2\text{O} = \text{H}_3\text{NO}_4$. British official acid $= 2\text{HNO}_3 + \text{H}_2\text{O}$, specific gravity 1.50, contains 85 per cent HNO_3 .

Nitric acid has sour taste; acid reaction with test papers; colors skin, feathers, wood, and albuminates in general a deep-yellow (picric acid); forms numerous organic compounds; removes H from H_2O , leaving NO_2 (nitro group) as substitute; oxidizes metals; dilute HNO_3 converts organic acids by oxidation into simpler acids; when H is substituted by metals it forms salts (nitrates).

Tests for HNO_3 : Cu, red vapors; indigo ($\text{C}_8\text{H}_5\text{NO}$) decolorized by conversion into isatin ($\text{C}_8\text{H}_5\text{NO}_2$); ferrous sulphate, brown liquid in presence of sulphuric acid; phenol (carbolic acid, $\text{C}_6\text{H}_5\text{OH}$) converted into yellow tri-nitro-phenol ($\text{C}_6\text{H}_2(\text{NO}_2)_3\text{OH}$); brucine, colored red; anilin sulphate in sulphuric acid, red color; diphenylamine ($\text{C}_6\text{H}_5\text{.NH.C}_6\text{H}_5$) in presence of sulphuric acid, deep blue.

Preparation of HNO_3 : Mostly from Chili saltpetre (NaNO_3) by distillation with sulphuric acid, $\text{NaNO}_3 + \text{H}_2\text{SO}_4 = \text{HNO}_3 + \text{NaHSO}_4$; the neutral salt, Na_2SO_4 , requires too high a heat so that HNO_3 would be wasted; 85 parts NaNO_3 + 98 parts H_2SO_4 = 63 parts HNO_3 ; the prod-

uct is red, and when converted into so-called *fuming nitric acid* it is rendered colorless by expelling NO_2 by current of air. *Red fuming nitric acid*, specific gravity 1.50, is obtained by distilling HNO_3 with strong sulphuric acid.

Nitrates artificially prepared by oxidation of nitrogenous substances in presence of alkaline base.

Nitric acid is *monobasic*: $\text{K}^{\text{I}}\text{NO}_3$; $\text{Na}^{\text{I}}\text{NO}_3$; $\text{Ag}^{\text{I}}\text{NO}_3$; $\text{Ca}^{\text{II}}(\text{NO}_3)_2$; $\text{Pb}^{\text{II}}(\text{NO}_3)_2$; $\text{Cu}^{\text{II}}(\text{NO}_3)_2$; $\text{Bi}^{\text{III}}(\text{NO}_3)_3$.

Aqua regia is a mixture of hydrochloric (3 volumes) and nitric acids (1 volume); it dissolves Au and Pt, has free Cl_2 and nitroxychloride (NO_2Cl) and nitrosylchloride (NOCl).

XV. CARBON.

Symbol, C. Tetrad. Atomic weight, 11.97. Specific gravity, 3.52 (diamond); 2.09 (graphite). Hypothetical vapor density, 11.97 ($\text{H} = 1$); 0.8316 (air = 1).

Occurs as diamond, graphite, anthracite, carbonic acid gas, carbonate of calcium, in organic tissues, etc.

Properties of C: Differ according to variety of the three allotropic forms—1, diamond; 2, graphite; 3, amorphous carbon (charcoal, etc.).

1. *Diamond*: Found in India, Borneo, Brazil, Capeland; either loose, in sand and detritus, or in itacolumite, xantophyllite. Crystals, transparent, regular system, rhombohedra, octahedra, hexakisdodekahedra, and hemihedric forms; mostly colorless, sometimes black (Bahia, 1848). Burns in oxygen (1694). Hardness = 10 (12), cuts glass when in hemihedric, hatchet-like form; only scratches are made with other facets or fractural curved crystals (?). Not altered by acids. Brittle; may be powdered in steel mortar. Highly refractive, index = 2.439. Cutting and polishing. Cleavage parallel to octahedron. Valuation by karat weight (3.2 troy grains). The following diamonds were found in central India:

Name.	Original weight. Karats.	Present weight. Karats.
Koh-i-Noor (England).....	900	102
Koh-i-Toor (Russia).....	193	193
Pitt, finest known (France).....	410	136

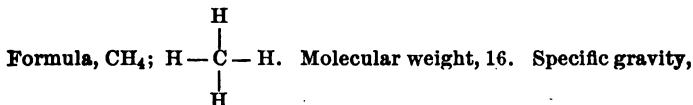
2. *Graphite* (plumbago, black lead): Occurs in the older strata, crystallized in hexagonal system; Burrowdale (England), Ceylon, Siberia. Used for pencils, lubricators, electric conductors, radiators. When heated with HNO_3 and chlorate of potassium forms graphitic acid; with permanganate of potassium, mellitic acid. Formed artificially in cast Fe. Burns with difficulty in O.

3. *Amorphous Carbon* (charcoal, lamp-black, anthracite, bone-black, gas carbon, coke, etc.): Burning charcoal, coking; contains mineral

substances (ashes) and hydrocarbons. Vegetable charcoal is a disinfectant, condenses gases, etc. Animal charcoal is a decolorizer. C at ordinary temperatures does not readily combine with other elements; at higher temperatures it readily unites with O and, therefore, is an active reductent.

Carbon and hydrogen: Large number of compounds are known; many occurring naturally in petroleum, others are products of distillation of organic bodies. Most of them to be treated in organic part of chemistry. Direct union of H with C by electric spark, acetylene, C_2H_2 . The simplest compound is marsh gas, CH_4 .

XVI. MARSH GAS.



7.985 (air = 1, specific gravity is 0.559). Synonyms, methane, methyl hydride, light carbureted hydrogen, fire damp.

Occurs in meadows, coal mines, and illuminating gas.

Properties: Odorless gas, nearly insoluble in H_2O . Burns with faintly luminous flame. Contains 75 per cent C and 25 per cent H. With two volumes O or one volume air forms explosive mixture: $CH_4 + 2O_2 = CO_2 + 2H_2O$. When heated to white heat CH_4 separates into C and $2H_2$.

Preparation: Heat equal quantities of sodium acetate and barium hydroxide in a retort; $2NaC_2H_3O_2 + Ba(OH)_2 = 2CH_4 + Na_2CO_3 + BaCO_3$.

Experiments: Explosion. Wire net. Davy lamp, etc.

XVII. ACETYLENE.

Formula, C_2H_2 ; $H - C \equiv C - H$. Molecular weight, 26. Specific gravity, 12.97 (air = 1, specific gravity is 0.898).

Colorless gas of peculiar, penetrating odor. Formed by direct union of C and H at high temperature. Present in coal gas and smoky flames. Burns with bright, luminous flame. Combines with Cu (also Ag) to form explosive compounds. Unites with H_2 to form C_2H_4 .

XVIII. OLEFIANT GAS.

Formula, C_2H_4 ; $H_2 = C = C = H_2$. Molecular weight, 28. Specific gravity, 18.97 (air = 1, specific gravity is 0.976). Synonyms, ethene, ethylene.

Properties: Colorless, irrespirable, odorous, poisonous gas. Contains 85.71 per cent C and 14.29 per cent H. Burns with bright flame. One volume forms explosive mixture with 3 volumes O or 15 volumes air. Cl unites with it to $C_2H_4Cl_2$; $C_2H_4 + Cl_2 = C_2H_4Cl_2$ (ethylene chloride).

Preparation: Distilling from a 1-litre flask 30 grammes 80 per cent alcohol with 30 grammes (130 cubic centimetres) sulphuric acid, avoiding heating of mixture, and adding 100 grammes pure sand; the gradual careful heating will decompose the alcohol; $C_2H_5O + H_2SO_4 = C_2H_4 + H_2SO_4 + H_2O$. The gas must be washed to remove sulphurous acid, carbon dioxide, and ether. As it dissolves in cold H_2O , it must be received over warm H_2O .

XIX. COAL GAS.

A mixture of various gases, according to the nature of the bituminous coal it is made from, contains:

	Per cent.
H.....	26 to 48
CH_4 and other paraffines	42 to 50
C_2H_4 " olefines	3 to 13
Carbon monoxide	7 to 8
C_2H_2	— to —

Illuminating power measured by comparison with candles of a burner using 5 cubic feet per hour.

Nature of flame.

Flame is gas in a state of high ignition. *Combustion* is the union of two or more elements accompanied with the phenomena of heat and light. *Incombustible* bodies may be highly heated so as to give red or white light, yet if they do not unite with other elements the term combustion does not apply to their state, but *glowing* or *incandescence*. Pure C heated in vacuum space (by electricity) becomes incandescent, but does not burn. H heated to enormously high temperatures in the *solar atmosphere* shows red flames, but does not unite with other elements and does not burn. Most substances used for producing artificial heat and light contain C and H, or either one. *Coal* burns with the atmospheric O to form one of the oxides of C; *hydrogen* burns forming H_2O . *Oil*, tallow, stearine, coal gas, wood, etc., contain both C and H, and are during combustion converted into gases, burning with flame. Bunsen burner, parts of flame (pages 31 and 32). Reducing temperature of flame, confining it by wire netting. Davy's safety lamp. Alternate burning of CH_4 in O, or of O in CH_4 .

Experiments: Davy's lamp. O flame in coal gas. Gunpowder in cool center of flame. Explosive mixture of coal gas with O.

XX. CARBON MONOXIDE.

Formula, CO; $C \equiv O$. Molecular weight, 28. Specific gravity, 13.966 (air = 1, specific gravity is 0.9674).

Properties: Colorless, inodorous, poisonous gas. Contains 42.86 per cent C and 57.14 per cent O. Condensable with difficulty. Little soluble in H_2O ($2\frac{1}{2}$ per cent by volume). Easily absorbed by solution of cuprous chloride in hydrochloric acid; burns with blue flame, forming, with O, carbon dioxide (CO_2). Reduces metallic oxides at high temperature. May be shown in blood spectrum. Unites in sunlight with Cl_2 to $COCl_2$, with Br_2 to $COBr_2$.

Preparation: (1) Heating mixture of oxalic acid with sulphuric acid and absorbing carbon dioxide by potassium hydrate; $C_2H_2O_4$ heated = $H_2O + CO + CO_2$. (2) Or 1 part potassium ferrocyanide is very carefully heated with 9 parts concentrated sulphuric acid and the gas passed through sodium hydrate; the reaction is very violent, and heat must be withdrawn as soon as it begins; $K_4Fe^{II}C_6N_6 + 6H_2SO_4 + 6H_2O = 6CO + 2K_2SO_4 + FeSO_4 + 3(NH_4)_2SO_4$. (3) Or carbon dioxide is passed through red-hot C; $CO_2 + \text{red-hot C} = 2CO$. (4) Or, as in the water gas process, steam is brought in contact with white-hot C; $H_2O + C = CO + H_2$.

Experiments: Burns with blue flame. Taper extinguished and re-kindled in CO. Absorption by cuprous chloride and re-evolution by heat.

XXI. CARBON DIOXIDE.

Formula, CO_2 ; $O = C = O$. Molecular weight, 44. Specific gravity, 21.948 (air = 1, specific gravity is 1.529). Synonyms, carbonic acid gas, choke damp.

Occurs in air (2.8 to 3.5 parts in 10,000 by volume), in most waters, liquid in quartz. Formed by combustion of C in air or O, or with oxides, by respiration of animals, putrefaction, fermentation, etc. Hence, its presence in mines and wells; Grotto del Canè, near Naples; Grottos at Eger, Pyrmont; Mofettos, near Vesuvius; Valley of Death, Java (crater of old volcano). In sparkling wines and mineral waters.

Preparation: By decomposition of carbonates by acids, $CaCO_3 + 2HCl = CaCl_2 + CO_2 + H_2O$; or by burning coal in air.

Properties: Colorless gas, with peculiar, pungent odor and taste. Acid reaction. Compressible to liquid by 36 atmospheres at $0^\circ C.$, 60 at 20° , or 73.5 at 30° . Liquid CO_2 is colorless, mobile, not miscible with H_2O , boils at $-78^\circ C.$ under 760 millimetres pressure, and by rapid evaporation becomes solid; when slowly evaporated, snow-like. Extinguishes flame. Irrespirable, poisonous, even when only a few per cent are mixed with air. Soluble in equal volumes of H_2O at $60^\circ C.$; at $0^\circ C.$ 1 volume H_2O dissolves about $1\frac{3}{4}$ volumes CO_2 . En-

tirely absorbed by caustic alkalis, lime, or baryta H_2O , forming carbonates. Heated metallic K or Na decomposes it; $\text{K}_4 + \text{CO}_2 = 2\text{K}_2\text{O} + \text{C}$. Heated Zn forms CO; $\text{Zn heated} + \text{CO}_2 = \text{CO} + \text{ZnO}$. Dark electric discharges form CO and O. H_2O charged with CO_2 dissolves carbonates of Ca, Ba, Fe, etc., which are otherwise insoluble. Hence hard H_2O from presence of carbonates. Aqueous solution gives acid reaction with color tests.

Experiments: Extinguishing candle. Pouring out of vessel (greater specific gravity than air). Absorption by caustic alkalies (chemical vacuum), bursting rubber cover of vessel. Precipitation and resolution of lime. Litmus reaction; boiling out.

Determination of C in organic substances by oxidation into CO₂, and absorbing this by caustic potash.

Human respiration in 24 hours produces 14,930 cubic inches (8,584 grains = $1\frac{1}{4}$ pounds) CO₂.

XXII. CARBONIC ACID.

Formula, H_2CO_3 ; $\text{O}=\text{C}=\begin{smallmatrix}\text{OH} \\ \text{OH}\end{smallmatrix}$. Dibasic.

Unknown in free state, as it decomposes at once into CO_2 and H_2O ; $\text{H}_2\text{CO}_3 = \text{H}_2\text{O} + \text{CO}_2$. Forms two series of salts by substitution of H_2 by metals. *Acid* carbonates or *bicarbonates*, when only one H is substituted (KHCO_3); *neutral* carbonates when both are substituted (K_2CO_3).

Cyanogen to be considered in organic part.

XXIII. CHLORINE GROUP.

Synonym, halogen elements.

Fluorine, F.....	= 18.98
Chlorine, Cl.....	= 35.37
Bromine, Br.....	= 79.77
Iodine, I.....	= 126.56

All four are monatomic. Unite directly with most elements. Form strong acids with H. At higher temperatures the stability decreases: HF , HCl , HBr , HI . F expels Cl; Cl expels Br; Br expels I. In their O acids the order is reversed and HIO_3 is most stable. No O compounds of F are known.

Fluorine.		Chlorine.		Bromine.		Iodine.	
Oxide.	Acid.	Oxide.	Acid.	Oxide.	Acid.	Oxide.	Acid.
.....	HF		HCl		HBr		HI
.....			HClO		HBrO		
.....			HClO ₂				
.....			HClO ₃				
.....			HClO ₄				
.....			HClO ₃		HBrO ₃		HO ₃
.....			HClO ₄		HBrO ₄		HO ₄

XXIV. CHLORINE.

Symbol, Cl. Monad (Cl¹). Atomic weight, 35·37567. Specific gravity, 35·37 (air = 1, specific gravity is 2·458). Specific gravity of liquid Cl, 1·38 (H₂O = 1). Discovered by Scheele, 1774. Supposed to be *oxy-murium*, a compound of hydrochloric acid with O. Proven to be an element by Gay-Lussac, 1809. Not free in nature. Principal compound, sodium chloride in rock salt and sea water.

Preparation: Heating hydrochloric acid with manganese dioxide; or a mixture of 5 parts sodium chloride and 5 parts manganese dioxide are mixed with a cooled solution of 12 parts sulphuric acid and 6 parts H₂O. When heated: $\text{MnO}_2 + 6\text{HCl} = \text{MnCl}_2 + 2\text{H}_2\text{O} + 2\text{Cl}_2$, or $\text{MnO}_2 + 2\text{NaCl} + 2\text{H}_2\text{SO}_4 = \text{MnSO}_4 + \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O} + \text{Cl}_2$. *Weldon's regeneration process* converts manganese chloride into calcium manganite (CaMnO₃?). *Deacon's process* makes impure chloride by passing a mixture of hydrochloric acid gas and air over hot bricks saturated with cuprous sulphate, and heated to 370° to 400° C. Cl dissolves in cold H₂O, so must be received over hot H₂O or by displacement. Is dried by passing over sulphuric acid or calcium chloride.

Properties: Yellowish-green gas. One litre = 3·1808 grains. Liquefies at -40° C., or by pressure of 4 atmospheres at 15° C. Liquid is greenish-yellow; not miscible with H₂O. Does not become solid at -90° C., boils at 33·6° C. Strong, pungent odor; very irritating to respiratory organs. Does not burn in air. Very soluble in H₂O (2·58 volumes in 1 of H₂O at 10° C.), forming Cl water. *Aqua chlori* (U. S. P.) contains 0·4 per cent Cl. A greenish liquid, easily decomposed by light into hydrochloric acid and H₂O. A crystalline Cl₂ + 10H₂O separates at 0° C. Has great affinity for other elements, forming *chlorides*, some uniting with appearance of light and heat. Affinity for H so great that it draws it from compounds, substituting H by Cl and forming HCl (hydrochloric acid). Liberates O as O₃ from H₂O; hence its bleaching and disinfecting properties. One atom O is replaced by 2 atoms Cl; $\text{CH}_4 + \text{Cl}_2 = \text{CH}_3\text{Cl} + \text{HCl}$; $3\text{H}_2\text{O} + 3\text{Cl}_2 = 6\text{HCl} + \text{O}_3$. Cl is absorbed by charcoal (200 volumes gas to 1 volume charcoal). Thin Cu foil, powdered As, or Sb placed in dry Cl unite in the dark, but with explosion in sunlight or other actinic rays.

Experiments: Bleaching moist calico. Burning Sb, P, thin Cu foil (Dutch leaf), or Na in Cl. Bringing flame of coal gas taper into jar of Cl. Igniting paper saturated with fresh oil of turpentine.

XXV. HYDROCHLORIC ACID.

Formula, HCl. Monobasic. Molecular weight, 36·375+. Specific gravity, 18·187 (air = 1, specific gravity is 1·264). Synonym, *muratic acid*.

Occurs in volcanic regions.

Formation: One volume Cl and 1 volume H unite (with explosive violence in a bright sunlight) to form 2 volumes HCl; also formed by action of Cl or H₂O, or other H compounds, or by acids on chlorides.

Preparation: In "Leblanc's soda process" immense quantities are formed by the action of sulphuric acid on sodium chloride; $2\text{NaCl} + \text{H}_2\text{SO}_4 = \text{Na}_2\text{SO}_4 + 2\text{HCl}$. Also by distilling 9 parts concentrated sulphuric acid, 5 parts sodium chloride, and $1\frac{1}{2}$ parts H₂O. Gas must be received over Hg on account of solubility in H₂O.

Properties: Colorless, pungent, irritating gas, fuming in air. Contains 97.25 per cent Cl and 2.75 per cent H. Compressible to colorless liquid by 40 atmospheres at 10° C. One volume H₂O absorbs 500 volumes HCl at 0° C., or 450 at 15°. This forms the liquid usually called hydrochloric acid. A saturated aqueous solution at 15° C. contains 40 per cent by weight of gas; specific gravity, 1.20, = 25° B. An acid containing 20.24 per cent HCl boils at 110° C. without change of composition; stronger or weaker is converted into this by boiling, and is composed of HCl + 8H₂O.

When acid is cooled to -21° or -22° C. and saturated with dry HCl gas, crystals form (HCl + 2H₂O), melting at -18° C. Some metals (Fe, Zn) decompose HCl, setting H free and forming chlorides, either when heated in dry gaseous HCl or when in aqueous solution. Some other metals (Cu, Hg, Au, Ag) do not decompose HCl. With metallic oxides HCl forms chlorides and H₂O; $\text{ZnO} + 2\text{HCl} = \text{ZnCl}_2 + \text{H}_2\text{O}$. When metallic salt solutions whose chlorides are insoluble (PbCl₂, Hg₂Cl₂, AgCl) are mixed with HCl, the chlorides are precipitated.

Acidum hydrochloricum (U. S. P.) = 31.9 per cent HCl + 68.1 per cent H₂O by weight; specific gravity, 1.16. 3.64 grammes = 31.9 cubic centimetres volumetric normal solution potassa.

Acidum hydrochloricum dilutum (U. S. P.) is made by mixing 6 parts stronger acid with 13 parts H₂O. Contains 10 per cent HCl; specific gravity, 1.049.

Acidum nitrohydrochloricum (U. S. P.), nitrohydrochloric acid, nitromuriatic acid, or aqua regia: Made by mixing 4 parts HNO₃ and 15 parts HCl. Dissolves Au. Contains Cl and two acids (page 105).

Nitroxyl chloride (NO₂Cl; $\text{O}=\text{N}=\text{O}$) is the result of direct union



of Cl with NO₂. A yellow liquid, boiling at 5° C. H₂O decomposes it into HNO₃ and HCl; $\text{NO}_2\text{Cl} + \text{H}_2\text{O} = \text{HNO}_3 + \text{HCl}$.

Nitrosyl chloride (NOCl; $\text{N}=\text{O}$) produced by union of NO (2 vol-



umes) and Cl (1 volume). A reddish-yellow gas, liquid below 0°C . H_2O decomposes it into HNO_2 and HCl ; $\text{NOCl} + \text{H}_2\text{O} = \text{HNO}_2 + \text{HCl}$.

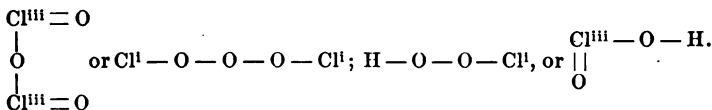
Chlorine monoxide (Cl_2O ; $\text{Cl}-\text{O}-\text{Cl}$) is prepared by passing Cl over HgO and cooling condensed product. A reddish-brown liquid. Boils at 21°C ., yellow vapor, very unstable, decomposes into O and Cl, explodes when heated. Dissolves in H_2O forming HClO ; $\text{H}_2\text{O} + \text{Cl}_2\text{O} = 2\text{HClO}$.

Hypochlorous acid (HClO ; $\text{Cl}-\text{O}-\text{H}$).

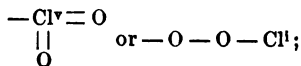
Prepared (1) by passing Cl over the hydroxide of alkalis or alkaline earths at cool temperatures. From these *hypochlorites* the acid may be obtained in aqueous solution. (2) Red mercuric oxide is stirred into H_2O and Cl passed, mercuric oxychloride and HClO are formed; $4\text{HgO} + \text{H}_2\text{O} + 2\text{Cl}_2 = \text{Hg}_4\text{O}_3\text{Cl}_2 + 2\text{HClO}$. (3) Calcium hypochlorite distilled with dilute HNO_3 .

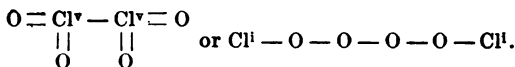
Properties: Colorless in dilute, yellow in concentrated solution. Decomposes slowly in dark, rapidly in light into Cl; H_2O ; HClO_3 . Very weak HClO does not expel CO_2 . Bleaches organic colors rapidly. Metallic Hg forms yellow oxychloride (soluble in HCl) with HClO .

Chlorine trioxide, or chlorous trioxide (Cl_2O_3), and *chlorous acid* (HClO_2): Their existence recently questioned. Some hold they are mixtures of other compounds. Cl_2O_3 formed by heating potassium chlorate with arsenous oxide and HNO_3 and cooling the greenish-yellow vapor to -20°C . Boils at 0°C .; explodes when heated; soluble in H_2O ; forms *chlorites* with alkali solutions. Dissected formulæ:



Chlorine tetroxide (Cl_2O_4 and ClO_2). *Formed* by treating potassium chlorate with concentrated sulphuric acid; also together with Cl by using HCl with KClO_3 . *Prepared* by mixing potassium chlorate (1 part), oxalic acid ($4\frac{1}{2}$ parts), sulphuric acid (1 part), H_2O (2 parts), and heating to 70°C . and condensing the gases. CO_2 escapes; Cl_2O_4 condenses to red liquid, boiling at 9.9°C . without decomposition; crystallizes at low temperature; soluble in cold H_2O . Vapor explodes on contact with organic substances. *Constitution*: Anhydride of chlorous and chloric acids; $\left. \begin{array}{l} \text{ClO} \\ \text{ClO}_2 \end{array} \right\} \text{O}$. Molecular weight, 67.5, but at low temperature Cl_2O_4 forms (analogous to NO_2 and N_2O_4); another example of dissociation. Dissected formula:

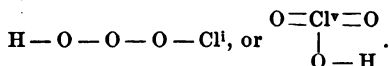




Experiments: Potassium chlorate heated with small quantity sulphuric acid forms ClO_2 , exploding by heat. HCl and potassium chlorate form Cl_2O_4 and Cl_2 . Burning P under H_2O at expense of O of Cl_2O_4 . Sugar, potassium chlorate, and sulphuric acid.

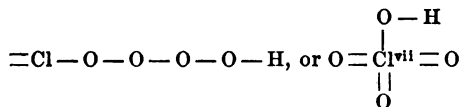
Chloric acid (HClO_3).

Obtained by adding dilute sulphuric acid to barium chlorate; $\text{Ba}(\text{ClO}_3)_2 + \text{H}_2\text{SO}_4 = 2\text{HClO}_3 + \text{BaSO}_4$. Decomposes in concentrated solutions; $4\text{HClO}_3 = 2\text{HClO}_4 + \text{Cl}_2 + \text{O}_2 + 2\text{H}_2\text{O}$. No corresponding oxide (Cl_2O_5). Formed by decomposition of HClO by daylight or heat. Salts formed by passing Cl into aqueous solutions of alkali hydroxides; $6\text{KOH} + 3\text{Cl}_2 = 5\text{HCl} + \text{KClO}_3 + 3\text{H}_2\text{O}$. Specific gravity of most concentrated solution, 1.282, $= \text{HClO}_3 + 7\text{H}_2\text{O}$; colorless and inodorous. Salts (the *chlorates*) are strong oxidizers, used in medicine and pyrotechnics, very explosive, caution in rubbing with S, P, C, etc. *Monobasic*. Dissected formula:



Experiments: Potassium chlorate and S in mortar. Potassium chlorate and antimony sulphide in cannon primers and friction matches. O by heating potassium chlorate (KClO_3); $\text{KCl}|\text{O}_3$; KCl shown by silver nitrate and O by glowing chip.

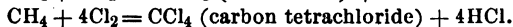
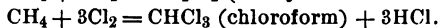
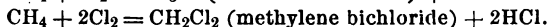
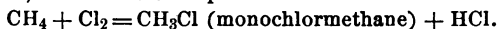
Perchloric acid (HClO_4): Most stable of O and Cl compounds. Formed by heating HClO_3 and chlorates. No corresponding oxide (Cl_2O_7). Heat decomposes KClO_4 into KCl and O_4 . Prepared by heating potassium chlorate (1 part) with sulphuric acid (4 parts) and distilling until the distillate ceases to solidify. This solid is impure HClO_4 , containing sulphuric acid. It is redistilled until crystals form in neck of retort. It is a colorless, corrosive, fuming liquid. Specific gravity, 1.782 at 15.5°C . Solidifies below 0°C . Melts at 15°C . Hygroscopic. Dissolves in H_2O with hissing sound; solutions very stable. Decomposes with explosion when concentrated. Hydrate, $\text{HClO}_4 + \text{H}_2\text{O}$, is crystalline. Dihydrate, $\text{HClO}_4 + 2\text{H}_2\text{O}$, is oily liquid, boils at 203°C ., distills without change. HClO_4 precipitates K, forming difficultly soluble crystalline KClO_4 . Dissected formula:



Nitrogen chloride (NCl_3): Formed by passing Cl through solution of sal ammoniac; $\text{NH}_4\text{Cl} + 3\text{Cl}_2 = \text{NCl}_3 + 4\text{HCl}$, or by electrolysis of sal

ammoniac at + pole. Is a dark-yellow, oily liquid. Specific gravity, 1.65; boils at 70° C.; insoluble in H₂O; explodes on contact with oily or fatty bodies, P, As, or Se, at ordinary temperatures with great violence; water of ammonia decomposes it. NCl₃ is NH₃ with the H₂ replaced by Cl₂.

Carbon chloride: Cl does not directly unite with C, but will by substitution take the place of H in *hydrocarbons*. These will be treated of in organic chemistry. Generally Cl₂ added to a hydrocarbon forms HCl, and leaves Cl in place of H:



XXVI. BROMINE.

Symbol, Br. Monad (Brⁱ). Atomic weight, 79.77. Specific gravity of vapor, 5.39 (air = 1); of liquid, 0° C., 3.1872. Boils, 63° C.

Occurs in combination with metals, especially Na, in sea water and saline springs (Kreuznach, Kanawha, W. Va.) and in some rock-salt deposits. Mother lye is distilled with sulphuric acid and manganese dioxide. Impure Br condenses and is purified from I and Cl.

Properties: Very dark red liquid; becomes solid at -7° C. Destroys organic tissue; vapor is irritating. Soluble in 80 parts H₂O. Below 4° C. forms crystalline hydrate, Br₂ + 10H₂O. Soluble in ether and carbon disulphide, which remove it from aqueous solution. Colors starch yellow. Expelled from compounds by Cl, but expels I. Resembles Cl in its chemical behavior. Great affinity for H. Monatomic in most compounds, but with O resembles Cl in variable higher atomicities. Its O acids are expelled by I, but expels those of Cl.

Hydrobromic acid (HBr).

Preparation: Can not be readily made, like HCl, by decomposing sodium salt with sulphuric acid unless when greatly diluted with H₂O. Sodium bromide, when so treated, furnishes free Br and sulphur dioxide. With phosphoric acid better results are obtained, but most HBr is made by decomposition of phosphorous tribromide by H₂O; PBr₃ + 3H₂O = 3HBr + H₃PO₃. Also formed by direct union of H and Br at high temperature. Concentrated aqueous solutions of HBr are made by mixing and, after reaction ceases, distilling 1 part red P, 10 parts Br, and 15 parts distilled H₂O.

Acidum hydrobromicum dilutum (U. S. P.) contains 10 per cent HBr. Specific gravity, 1.077. May be made in such dilute solution from potassium bromide and dilute sulphuric acid. 16.2 grammes = 20 cubic

centimetres normal volumetric solution of soda. Specific gravity of 50-per-cent HBr = 1.513.

Description: Molecular weight, 80.77. A gas closely resembling HCl; specific gravity, 2.79; very soluble in H₂O; condensed to liquid by pressure or at -70° C., crystallizes at lower temperatures. Boiling leaves a 48-per-cent acid (HBr + 5H₂O); boils at 125°; specific gravity, 1.49 at 14° C. Specific gravity of saturated aqueous solution at 0° C. = 1.78; by cooling forms crystals of HBr + 2H₂O, melting at -11° C. Exposed to air and light aqueous HBr becomes red from free Br. Cl liberates Br from HBr and bromides.

Oxygen and bromine: No oxides but corresponding acids; hypobromous (HBrO), bromic (HBrO₃), and perbromic (HBrO₄) acids.

Hypobromous acid (HBrO): Formed by action of Br on mercuric oxide in H₂O; $\text{HgO} + 2\text{Br}_2 + \text{H}_2\text{O} = \text{HgBr}_2 + 2\text{HBrO}$. Alkaline solution of sodium hypobromite used in Yron's and in Esbach's methods of determining urea in urine, by converting it into N and CO₂ and H₂O; $\text{CON}_2\text{H}_4 + 3\text{NaBrO} + 2\text{NaOH} = \text{N}_2 + 3\text{NaBr} + \text{Na}_2\text{CO}_3 + 3\text{H}_2\text{O}$. Solution made of 17 grammes caustic soda, 5 grammes Br, and 133 cubic centimetres H₂O.

Bromic acid (HBrO₃): Formed by passing Cl in Br water; $\text{Br}_2 + 6\text{H}_2\text{O} + 5\text{Cl}_2 = 10\text{HCl} + 2\text{HBrO}_3$. Adding Br to hot aqueous solutions of caustic alkalies or baryta, forms bromates; $3\text{Br}_2 + 6\text{NaOH} = 5\text{NaBr} + \text{NaBrO}_3 + 3\text{H}_2\text{O}$. Sulphuric acid liberates aqueous solution of HBrO₃ from barium bromate. Free Br liberates dilute HBrO₃ from silver bromate. HBrO₃ can not be concentrated beyond HBrO₃ + 7H₂O without decomposition. When heated it forms Br and O and H₂O; $2\text{HBrO}_3 = \text{Br}_2 + \text{O}_2 + \text{H}_2\text{O}$. *Properties* are similar to HClO₃ which it expels from bases.

Perbromic acid (HBrO₄): Formed by action of Br on HClO₄; $2\text{HClO}_4 + \text{Br}_2 = 2\text{HBrO}_4 + \text{Cl}_2$. *Properties*, similar to HClO₄. Little known.

Bromine chloride (BrCl): Formed by direct union of Br and Cl at 0° C. Red-brown liquid, mobile, boils at 13° C., and decomposes into Cl and Br. Forms crystalline hydrate.

Nitrogen bromide (NBr₃): Formed from NCl₃ and potassium bromide; $\text{NCl}_3 + 3\text{KBr} = \text{NBr}_3 + 3\text{KCl}$. Explosive. Little studied.

XXVII. IODINE.

Symbol, I. Monad (I'). Atomic weight, 126.56. Specific gravity of solid, 4.958; of vapor (air = 1), 8.85. Molecular weight up to 600° C., 253.12; above 1,400° C., 126.56 (dissociated). Boils above 200° C. Not free in nature. Found mostly as iodide of Na, Mg, etc., in sea water, marine plants, and animals. Also in Chili saltpeter, and as iodide of Pb and of Ag. Made from ashes of kelp. These are lixivi-

ated and the iodides remain in last liquid. These are distilled with sulphuric acid and manganese dioxide, or with Cl or HNO_3 , and the raw I purified by sublimation. Discovered by Courtois, of Paris, 1812.

Properties: Black or gray solid rhombic crystals. Melts at 114°C . Deep violet-blue vapor. Specific gravity of vapor at different temperatures proves dissociation of I_2 above $1,400^\circ \text{C}$. into 2I . Under reduced pressure dissociation begins at 400°C . At 10°C . 1 part soluble in 5,500 of H_2O ; at 15°C . 1 part in 11 of alcohol. Freely soluble in ether, glycerin, aqueous solution of potassium iodide, carbon disulphide, and chloroform. Magenta color in two latter, brown in other solvents. Colors starch blue-black (few exceptions). No hydrate known. Less affinity for H than Cl or Br. Expelled from iodides by Cl and Br. First employed in medicine by Dr. Coindet, of Geneva, 1819. Tincture, United States Pharmacopœia, 1880 (I, 8: alcohol, 92). Compound tincture, United States Pharmacopœia, 1870 (I, 1: KI, 2: alcohol, 16). Colorless tincture by digesting with NH_3 , or instantaneously by NH_3 and phenol; triethylammonium iodide formed. I corrodes skin, brown color; poisonous; antidote is starch. Forms *iodides*.

Hydriodic acid (HI): Aqueous solution made by passing sulphureted H through I suspended in H_2O , and distilling at 127°C .; a 57-per-cent acid distills. When saturated with I (10 parts I and 1 of red P to 5 parts of solution) and heated, HI gas evolves.

Properties: Colorless, fuming gas, soluble in H_2O . Specific gravity (air = 1), 4.42. Condensed at 0°C . by four atmospheres. Solid at -51°C . Decomposed by light or heat. Boiled aqueous solution = 57.75 per cent HI ($\text{HI} + 5\text{H}_2\text{O}$); boils at 127°C .; specific gravity, 1.67. When saturated at 0°C . specific gravity = 2.0. Solution is colorless, but soon becomes brown from free I. Many oxidizing substances, Cl and Br, set I free from iodides and HI. At high temperatures HI is a reductant.

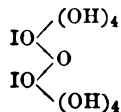
Oxygen compounds of I: Iodine pentoxide (I_2O_5) and iodine peroxide (I_2O_7); iodic acid (HIO_3) and periodic acid (HIO_4).

Iodic acid (HIO_3): Forms *iodates*. Prepared by dissolving I in concentrated HNO_3 , or by passing Cl through 1 part I in 20 of H_2O ; specific gravity, 1.50. Colorless, rhombic prisms; specific gravity, 4.629. Melts at -15°C ., at -17°C . forms hexagonal plates = $(\text{HIO}_3)_2 + 9\text{H}_2\text{O}$. More stable than HClO_3 or HBrO_3 . Unchanged at 100°C .; at 170°C . it loses H_2O and becomes I_2O_5 ; $2\text{HIO}_3 = \text{I}_2\text{O}_5 + \text{H}_2\text{O}$.

Iodine pentoxide (I_2O_5): Made by union of O_3 and I vapor. White, crystalline powder; specific gravity, 4.48. At 300°C . decomposes into I and O.

Periodic acid (HIO_4): Formed by action of I on HClO_4 . Best isolated from Ag salt (AgIO_4). Rhombohedral crystals = $\text{HIO}_4 + 2\text{H}_2\text{O}$,

soluble in H_2O . At 200°C . decomposes into I_2O_7 and H_2O , afterward into I_2O_5 ; O and H_2O . Salts exist which suggest condensed polyiodic acids; diiodic acid:



Chlorine and iodine: Iodine monochloride (ICl); iodine trichloride (ICl_3); and possibly iodine tetrachloride (ICl_4).

Iodine monochloride (ICl): Prismatic, hyacinth-red crystals; fuse at 24.7°C .; boil at 100.5°C .; decompose by H_2O to HCl and HIO_3 , slowly changed to I and ICl_3 and $\text{ICl}_4(?)$ in red octahedra crystals. Made by passing dry Cl over I , dissolving I in nitrohydrochloric acid, and extracting ICl from solution with ether.

Iodine trichloride (ICl_3): Yellow crystals, sublime at 15°C .; easily dissociated in air at 25°C .

Iodine monobromide (IBr): Hard solid; melts at 36°C .; sublimes, color like I . Made by direct union of I and Br .

Nitrogen and iodine (NH_2I ; NHI_2 and NI_3): Very explosive, dark-brown solid. Made by replacing H in NH_3 , or by digesting I in strong NH_3 .

XXVIII. FLUORINE.

Symbol, F . Monad (F^1). Atomic weight, 18.98.

Has great affinity for all elements, *except* O , and can not be isolated in a vessel without attacking it. Hence its properties are difficult to study. Occurs as fluorspar (CaF_2), cryolite ($\text{Al}_2\text{F}_6 \cdot 6\text{NaF}$), in many minerals, enamel of teeth, bones, etc. Most probably a gas. Great tendency to unite with Si and with H .

Hydrofluoric acid (HF): Colorless liquid; specific gravity, 0.9879 at 12°C .; boils at 19°C ., remains liquid at -34°C . Easily soluble in H_2O . Colorless, fuming, poisonous vapor. Made by heating fluorspar or cryolite with sulphuric acid. Aqueous solution boiled at 120°C . forms a permanent 38-per-cent acid ($\text{HF} + 2\text{H}_2\text{O}$). Corrosive. Dissolves most metals, forming *fluorides*. Etches glass, forming silicon tetrafluoride, SiF_4 . Sold in gutta-percha bottles. Schwankhard, of Müsenburg, etched glass, 1670. Scheele showed this to be due to HF , 1771. Gay-Lussac and Thenard studied HF . Prat studied pure F .



Cuprammonium oxyfluorides: $\text{CuOHF}(\text{NH}_2)_2$; $\text{CuOHF}(\text{NH}_3)_2$.

XXIX. SULPHUR GROUP.*

Oxygen, O	= 15·9633	Selenium, Se	= 78·7976
Sulphur, S	= 31·9840	Tellurium, Te	= 127·9600

All diatomic, but also in some compounds tetratomic (also hexatomic). Do not unite with H at ordinary temperature, incompletely at higher temperature, but readily in *statu nascente*: H_2S , H_2Se , H_2Te . When heated with O they form *dioxides*: SO_2 , SeO_2 , TeO_2 . *Diatomic* to H and metals. *Diatomic* and *tetratomic* to Cl group. *Hexatomic* and *tetratomic* to O. Scale of affinities to H: O (greatest), S, Se, Te (least). Scale of affinities to O: S (greatest), Se, Te (least). Sulphur dioxide (SO_2) in presence of H_2O reduces SeO_2 to Se and H_2SO_4 (sulphuric acid); Te_2O to Te and H_2SO_4 . H_2 with O = 57,200 calories; with S, 4,600; with Se, —5,400.

XXX. SULPHUR.

Symbol, S. Diad and hexad (S^{II} and vi). Atomic weight, 31·9840. Specific gravity of solid S: natural rhombic, 2·07; melted prismatic, 1·98; amorphous, 1·95. Specific gravity of vapor (air = 1) at 500°C ., 6·65 ($\text{H} = 1$, 96); at $1,000^\circ\text{C}$., 2·216 ($\text{H} = 1$, 31·984).

Occurs free in volcanic regions of Girgenti, Sicily; Utah; California; Urbino and Reggio, Italy. Found at Rodoboy, Crotia, etc., with limestone, gypsum, marl, etc.; as sulphides (pyrites, blende, glance, etc.); widely distributed as sulphate (gypsum); also in plants and animals.

Prepared by melting or distilling the natural minerals, or by subliming from pyrites. *Purified* by resublimation. Found in commerce as flower of sulphur, sublimed or brimstone (in mass or cast in sticks called roll sulphur). About 80,000 tons annually from Sicily. S from pyrites often impure (As, Tl).

Physical properties: Yellow, becoming colorless at low temperature (-50°C .); brittle, becomes electric (—) by friction. Three varieties: 1. *Natural*, or crystallized at low temperatures from carbon disulphide. Yellow, transparent, rhombic crystals. 2. Monoclinic *crystals* from melted S; transparent, but soon become opaque by splitting into small rhombic octohedra. Both these forms soluble in carbon disulphide (46 parts in 100), hydrocarbons, oils, sulphur monochloride; little soluble in alcohol or ether. S fuses at $111\cdot5^\circ$ to 118°C . to thin, yellow, mobile liquid; higher heat (195°C .) makes it darker and viscid; up to 230°C . it will not flow, above that it becomes thinner and boils at 447°C .; yellow vapor. 3. When suddenly chilled at brown stage (230°C .) it becomes *amorphous*, elastic like rubber, but

*Oxygen group.

afterwards brittle again. Partially soluble in carbon disulphide; leaves yellow, insoluble residue (similar in sublimed S); at 100° C. it is reconverted into ordinary S. *Precipitated S* made by adding HCl to calcium sulphide. S vapor at 500° C. = S₈; at higher temperature this dissociates, beginning at 700° C., complete at 1,000° C., when the molecules = S₂. Crystallized S is electro-negative; amorphous is electro-positive.

Chemical properties: Analogous to O—H₂O and H₂S, CO₂ and CS₂, KOH and KSH, K₂O and K₂S, HgO and HgS. Metals combine directly with S; some burn in its vapor (Ca, Fe). With O at 260° C. it burns, forming SO₂. S forms *sulphides*.

XXXI. HYDROGEN SULPHIDE.

Formula, H₂S. Molecular weight, 34. Specific gravity, 17 (air = 1; specific gravity, 1.178). Synonyms: Hydrogen sulphide, sulph-hydric acid, hydrothionic acid.

Occurs in volcanic regions and mineral waters. *Formed* by putrefaction of organic substances containing S, also by decomposition of sulphides, and by union of S and H in nascent state, or by passing steam over boiling S in closed tubes at 200° C. *Usually prepared* by ferrous sulphide and dilute sulphuric acid, or calcium sulphide and HCl. Paraffin and S.

Properties: Colorless, odorous gas; specific gravity, 1.178 (air = 1); condensed by 15 atmospheres pressure at 10° C., or ordinary pressure at -74° C., to colorless liquid; specific gravity, 0.9. Solidifies at -85° to white crystals. Soluble in H₂O; at 0° C., 4.3 in 1 of H₂O; at 15°, 8.2 in 1. Solution rapidly oxidizes to H₂O and S. When heated to 350° C. to 400° C. it dissociates. HNO₃ and other oxidizers decompose it. Cl, Br, and I form S and HCl, HBr, and HI (I only in presence of H₂O). Strong reducent. Metals heated in H₂S decompose it and form sulphides. Many metallic salts are converted into sulphides and are precipitated. Some from acid solution (As, Cd, yellow; Sb, orange; Sn, brown; Ag, Au, Pb, Bi, black); and some from alkaline solution (Fe, Ni, black; Zn, white; Co, blue; Mn, flesh color). Some are not precipitated (Mg, Ca, Ba, Sr, K, Na, Li). Hence, H₂S is a valuable analytical reagent. Pb salts are very sensitive to it.

Experiments: Metallic precipitates As, Cd, Sb, Au, Fe, and Zn. Hare's experiment. H₂S is a weak acid. Burns in air to SO₂ and H₂O.

Hydrogen disulphide (H₂S₂, persulphide): Yellow, odorous oil. Formed when calcium sulphide is treated with HCl. Dissolves S. Decomposes at 60° C. into S and H₂S. Forms crystalline compounds with strychnine.

Polysulphides: Possibly the series H_2S_2 , H_2S_3 , and H_2S_6 , similar to CaS_2 , CaS_3 , and CaS_6 .

XXXII. OXYGEN COMPOUNDS OF SULPHUR.

Oxides of sulphur: Dioxide, SO_2 ; trioxide, SO_3 ; dithionic oxide, S_2O_3 ; heptoxide, S_2O_7 .

Sulphur acids:

H_2S , sulphydric.	$H_2S_2O_3$, thiosulphuric (hyposulphurous, dithionous).
H_2SO_2 , hydrosulphurous (hyposulphurous).	$H_2S_2O_6$, dithionic (hyposulphuric).
H_2SO_3 , sulphurous (monothionous).	$H_2S_3O_6$, trithionic.
H_2SO_4 , sulphuric.	$H_2S_4O_6$, tetrathionic.
$H_2S_2O_7$, pyrosulphuric (Nordhausen).	$H_2S_5O_6$, pentathionic.

Sulphur Dioxide:

Formula, SO_2 . Molecular weight, 63·910+. Specific gravity, 63·9066 (air = 1, specific gravity is 2·24); of liquid, 1·49.

Occurs in volcanic regions. *Formed* by burning S in O; heating metallic oxides with S; roasting metallic oxides; heating some sulphates; reducing H_2SO_4 with S, C, metals, etc.

Prepared by heating H_2SO_4 with Cu.

Properties: Colorless gas, stifling odor; liquid at $-18^\circ C.$, or by 2 to 3 atmospheres at ordinary temperature; boils at $-8^\circ C.$, crystallizes at $-76^\circ C.$; absorbed by H_2O , forming dilute H_2SO_3 . *Concentrated solution* contains 9·5 per cent SO_2 ; specific gravity, 1·046; heat dispels the SO_2 . Saturated aqueous solution at $0^\circ C.$ forms crystals of $SO_2 + 15H_2O$ or $H_2SO_3 + 14H_2O$; these melt at $4^\circ C.$ and separate into SO_2 and H_2O . SO_2 oxidizes in contact with H_2O to H_2SO_4 by O_3 , Pt black, or slowly in air. SO_2 is also converted in H_2SO_4 by Cl, Br, or I in aqueous solution, HCl, HBr, or HI being formed. By its tendency to unite with O and H_2O to form H_2SO_4 , it is a great reductent of metallic oxides, chromic acid, HNO_3 , permanganic acid, etc.; also bleaches vegetable colors, straw, etc. An aqueous solution heated in closed tube to 170° to $180^\circ C.$ decomposes into S, H_2S , and H_2SO_4 ; $6H_2O + 7SO_2 = S + H_2S + 5H_2SO_4$. Used as a disinfectant. Salts of $(H_2)SO_3$ (*sulphites*) form in two series: *acid* ($NaHSO_3$) and *neutral* (Na_2SO_3).

Chlorine compounds of sulphur oxides.

Sulphuryl chloride (SO_2Cl_2): Formed by SO_2 with dry Cl_2 in sunlight. Colorless, fuming, odorous liquid. Specific gravity, 1·66. Boils at $77^\circ C.$ Decomposed by H_2O to $H_2SO_4 + 2HCl$.

Thionyl chloride ($SOCl_2$): Formed by action of phosphorous pentachloride on SO_2 ; $PCl_5 + SO_2 = POCl_3 + SOCl_2$. Colorless, pungent liquid. Boils at $78^\circ C.$ Decomposed by H_2O into H_2SO_3 and $2HCl$.

Sulphuryl oxychloride; chloranhydride of sulphuric acid ($\text{SO}_2\cdot\text{OH}\cdot\text{Cl}$): Derived from $\text{SO}_2\cdot\text{OH}\cdot\text{OH}$ by action of phosphorous pentchloride, also found by direct union of SO_3 and HCl . Colorless, fuming liquid. Specific gravity is 1.776 at 18°C . Boils at 152°C . Potassium chlorosulphate ($\text{SO}_2\cdot\text{Cl}\cdot\text{OK}$).

Pyrosulphuryl chloride, or chloranhydride of disulphuric acid ($\text{S}_2\text{O}_5\text{Cl}_2$): Boils at 146°C .

Hydrosulphurous, or hyposulphurous acid $\left\{ \text{H}_2\text{SO}_2; \text{SO} \begin{smallmatrix} \text{OH} \\ \text{H} \end{smallmatrix} \right\}$: Monobasic ($\text{HO}\cdot\text{SHO}$); $\text{NaO}\cdot\text{SHO}$. Formed by action of Zn on aqueous solution of SO_2 ; $2\text{SO}_2 + \text{Zn} + \text{H}_2\text{O} = \text{ZnSO}_3 + \text{H}_2\text{SO}_2$; also by action of nascent H . Yellow, unstable solution. Reductant. Bleaches more rapidly than SO_2 . Salts (*hydrosulphites*) are stable.

XXXIII. SULPHURIC ACID.

Formula, H_2SO_4 ; $\text{SO}_2 \begin{smallmatrix} \text{OH} \\ \text{OH} \end{smallmatrix}$. Dibasic. Molecular weight, 97.2372.

Synonym, oil of vitriol. Specific gravity, 1.854.

Occurs free in volcanic lakes; as sulphate in gypsum, copperas, zinc vitriol, etc. Formed by oxidation of S by HNO_3 , or of SO_2 by O and H_2O . Manufactured on commercial scale in Pb chambers, where SO_2 is formed by combustion of S or pyrites and comes in contact with HNO_3 : (1) $3\text{SO}_2 + 2\text{HNO}_3 + 2\text{H}_2\text{O} = 3\text{H}_2\text{SO}_4 + 2\text{NO}$; (2) $2\text{NO} + 2\text{O} = 2\text{NO}_2$; (3) $2\text{NO}_2 + 2\text{H}_2\text{O} + 2\text{SO}_2 = 2\text{H}_2\text{SO}_4 + 2\text{NO}$, and begins at 2 again, or $3\text{NO}_2 + \text{H}_2\text{O} = 2\text{HNO}_3 + \text{NO}$, and begins at 1 again. Some N_2O is also formed, which, with the residuary gases, is condensed in Gay-Lussac's tower, filled with coke and H_2SO_4 , which absorbs the N oxides. The latter are mixed with the hot SO_2 in a Glover tower, interposed between Pb chambers and combustion furnace. If

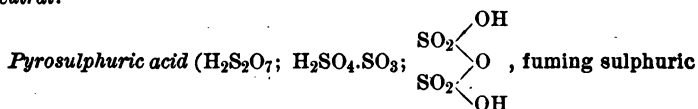
H_2O be deficient, the crystals of the Pb chamber form $\text{SO}_2 \begin{smallmatrix} \text{NO}_2 \\ \text{OH} \end{smallmatrix}$, which

decomposes with H_2O ; $\text{SO}_2 \begin{smallmatrix} \text{O}-\text{NO} \\ \text{O}-\text{H} \end{smallmatrix} + \text{H}_2\text{O} = \text{SO}_2 \begin{smallmatrix} \text{OH} \\ \text{OH} \end{smallmatrix} + \text{NO}\cdot\text{OH}$, it

dissolves in H_2SO_4 . Acid of the chambers has specific gravity, 1.50; contains 60 per cent H_2SO_4 ; is concentrated in Pb pans to 80 per cent acid; specific gravity, 1.72. Further concentration in Pt vessels to 92 per cent acid; specific gravity, 1.83. This is crude oil of vitriol. When distilled, first one-third is aqueous acid, then at 330°C . nearly pure H_2SO_4 (specific gravity, 1.854 at 0°C . or 1.842 at 12°C .), containing 1.5 per cent H_2O . Latter forms white crystals at -35°C .; puri-

fied by recrystallization; melt at $10.5^{\circ}\text{C}.$; at 40° white fumes of SO_3 are formed.

Dissociation of H_2SO_4 by heat; $\text{H}_2\text{SO}_4 = \text{SO}_3 + \text{H}_2\text{O}$, which unite again on cooling. $\text{H}_2\text{SO}_4 + 1\text{H}_2\text{O}$ cooled to $0^{\circ}\text{C}.$ crystallizes, melting at $7.5^{\circ}\text{C}.$; boil at $205^{\circ}\text{C}.$, losing H_2O . Attacks H_2O with great avidity, producing heat. Danger of mixing large quantities of H_2SO_4 and H_2O . Soluble in H_2O in all proportions. $\text{H}_2\text{SO}_4 + 2\text{H}_2\text{O}$ contracts to maximum density; specific gravity, 1.62; boils away H_2O at $162^{\circ}\text{C}.$ H_2SO_4 passed over red-hot bricks, etc., splits into $\text{SO}_2 + \text{H}_2\text{O} + \text{O}$, hence used for O manufacture. Attacks nearly all metals, except Pt and Pb, forming *sulphates*. Separates most other acids from their salts. Ba salt is insoluble. Forms two series of salts: *acid* and *neutral*.



acid, Nordhausen sulphuric acid): Made from ferrous sulphate at Nordhausen and in Bohemia; $6\text{FeSO}_4 + 3\text{O}$ (copperas roasted) $= 2\text{Fe}_2(\text{SO}_4)_3 + \text{Fe}_2\text{O}_3$ and $\text{Fe}_2(\text{SO}_4)_3 = \text{Fe}_2\text{O}_3 + 3\text{SO}_3$, or if water of crystallization is not entirely expelled from copperas, $6\text{FeSO}_4 + 2\text{H}_2\text{O} + 2\text{O}_2 = 2\text{Fe}(\text{SO}_4)_2(\text{OH})_2 + \text{Fe}_2\text{O}_3$. Is in large, transparent, colorless crystals, melting at $35^{\circ}\text{C}.$ Heat decomposes it into H_2SO_4 and SO_3 .

Sulphur trioxide (SO_3 , sulphuric acid anhydride): Formed directly from SO_2 and O by dark electric discharge, or contact with Pt sponge. Best made by careful distillation of $\text{H}_2\text{S}_2\text{O}_7$. Is in large, transparent prisms. Melts at $14.8^{\circ}\text{C}.$ Boils at $46.2^{\circ}\text{C}.$ Fumes in air. Attacks H_2O strongly, and hisses when thrown into it, forming H_2SO_4 . Passed through red-hot tube it dissociates into SO_2 and O. When containing H_2SO_4 it forms silky, asbestos-like felt, unchanged at $50^{\circ}\text{C}.$, but melted at higher temperature.

Dithionic oxide (S_2O_3 , sulphur sequioxide): When powdered S is added to melted SO_3 , blue drops form, which remain on evaporating the excess of SO_3 , and form crystals, bluish-green when dry, easily decomposed into S and SO_2 . Deliquesces. Forms S; SO_3 ; H_2SO_4 and *polythionic acid* with H_2O .

Sulphur heptoxide (S_2O_7 , persulphuric anhydride): Made by dark electric discharges through $2\text{SO}_2 + \text{O}_4$ (equal volume) $= \text{S}_2\text{O}_7 + \text{O}$, and cooling, or through 2SO_3 and excess of O, and cooling. Liquid; solidifies at $0^{\circ}\text{C}.$ to long transparent needles. Only stable below $0^{\circ}\text{C}.$; dissociates above $0^{\circ}\text{C}.$ Decomposed by H_2O . Dissolves in H_2SO_4 . Also formed by electrolysis of H_2SO_4 , or by H_2O_2 and H_2SO_4 .

Dithionous acid ($\text{H}_2\text{S}_2\text{O}_3$, hyposulphurous or thiosulphuric acid):

Only known in salts. When liberated it dissociates $\left\{ \text{SO}_2 \begin{matrix} \text{SH} \\ \text{OH} \end{matrix} \right\}$ into $\text{SO}_2 + \text{S} + \text{H}_2\text{O}$. More stable in great dilution. Na salt made by boiling neutral sodium sulphite with S; $\text{Na}_2\text{SO}_3 + \text{S} = \text{Na}_2\text{S}_2\text{O}_3$. Much used in photography.

Dithionic acid ($\text{H}_2\text{S}_2\text{O}_6$; $\begin{matrix} \text{SO}_2-\text{OH} \\ | \\ \text{SO}_2-\text{OH} \end{matrix}$), hyposulphuric acid, one of the

polythionic series): Aqueous solution made by decomposing its Ba salt with H_2SO_4 . May be concentrated to specific gravity, 1.347; above this it decomposes into SO_2 and H_2SO_4 .

Trithionic acid ($\text{H}_2\text{S}_3\text{O}_6$): Only known in salts. Made by digesting potassium acid sulphite with S at 50° to 60° C.; $6\text{KHSO}_3 + 2\text{S} = 2\text{K}_2\text{S}_3\text{O}_6 + \text{K}_2\text{S}_2\text{O}_8 + 3\text{H}_2\text{O}$; or by evaporating a solution of KHSO_3 and KHS_2O_3 .

Tetrathionic acid ($\text{H}_2\text{S}_4\text{O}_6$): Only known in salts. Formed by action of I on hyposulphites; $2\text{K}_2\text{S}_2\text{O}_3 + \text{I}_2 = 2\text{KI} + \text{K}_2\text{S}_4\text{O}_6$. Is a constant result in volumetric determinations by normal I solution.

Pentathionic acid ($\text{H}_2\text{S}_5\text{O}_6$): Unknown in concentrated solutions; only in aqueous solution. Made by passing H_2S through watery solutions of SO_2 ; $5\text{SO}_2 + 5\text{H}_2\text{S} = \text{H}_2\text{S}_5\text{O}_6 + 5\text{S} + 4\text{H}_2\text{O}$. The *polythionous acids* are: $\text{H}_2\text{S}_2\text{O}_3$; $\text{H}_2\text{S}_2\text{O}_6$; $\text{H}_2\text{S}_3\text{O}_6$; $\text{H}_2\text{S}_4\text{O}_6$, and $\text{H}_2\text{S}_5\text{O}_6$. Their Ba salts are soluble. All decompose from concentrated solutions, by boiling, into S, H_2SO_4 , and SO_2 .

Chlorides of sulphur: Monochloride, S_2Cl_2 ; dichloride, SCl_2 ; tetrachloride, SCl_4 .

Sulphur monochloride (S_2Cl_2): Formed by passing Cl over melted S or heated powdered S. Reddish-yellow liquid; irritating to the eyes. Specific gravity, 1.68 to 1.709 at 0° C.; boils at 139° C. Fumes in air; decomposes with H_2O into HCl, S, and SO_2 . Decomposed by Sb, As, Sn, and Al, but not by Na or Mg. Readily dissolves S. Used in vulcanizing rubber.

Sulphur dichloride (SCl_2): Made by saturating S_2Cl_2 with Cl_2 at 0° C. Dark-red liquid; specific gravity, 1.62; boils at 64° C., decomposing into S_2Cl_2 and Cl_2 . Decomposed by H_2O . Dissociates partially at common temperatures, fully at 130° C.

Sulphur tetrachloride (SCl_4): Formed by saturating S_2Cl_2 with Cl_2 at -20° to -22° C. Yellowish-brown, mobile liquid. Dissociates above -20° C., completely at 6° C. Forms crystals with SnCl_4 .

Bromides and iodides of sulphur: No pure Br compounds isolated. S dissolves in Br, forming S_2Br_2 ; smell like S_2Cl_2 ; boils at 190° C. The official iodide of sulphur, S_2I_2 , made by heating S with I.

Carbon and sulphur: Carbon disulphide, CS_2 . Carbon monosulphide, CS. Sulphocarbonic acid, H_2CS_3 . Carbon oxysulphide, COS.

Carbon disulphide (CS_2): Made by adding S to red-hot charcoal in Fe cylinders, and cooling distillate. Colorless liquid, highly refractive (index 1.70); ethereal odor when pure, fetid when impure; boils at 1.292 ; vapor density, 2.67 . Insoluble in H_2O ; dissolves in alcohol; great solvent of S, I, P, fats, and resins. Vapor mixed with O is explosive. Nascent H with CS_2 forms CH_2S crystals, insoluble in H_2O , alcohol or ether; soluble in CS_2 . Vapor easily dissociated by heat; when passed over metallic oxides forms metallic sulphides and CO_2 ; at temperatures below -3°C . it forms a crystalline hydrate.

Sulphocarbonic acid (H_2CS_3): The salts are formed by CH_2S with alkaline sulphides, from which HCl separates the acid as a red-brown oil.

Carbon monosulphide (CS?): Formed by decomposing CS_2 exposed to sunlight in sealed tubes. A chestnut-brown, tasteless, odorless, little soluble powder; specific gravity, 1.66 . Easily decomposed into C and S.

Carbon oxysulphide (COS): An aromatic gas found in mineral springs. Specific gravity, 2.106 . Absorbed by caustic potash solutions. Made by passing CO and S through red-hot tubes.

Sulphur and nitrogen: S does not unite directly with N, but ammonia acts on sulphur chlorides and oxychlorides. NS is a yellow powder, insoluble in H_2O , soluble in CS_2 , crystallizing from it in transparent orange crystals; sublimes at 135°C ., melts at 158°C ., explodes by percussion or by heating to 207°C . Alkaline solutions decompose it into NH_3 , sulphites, and hyposulphites. When treated with Cl while suspended in chloroform it dissolves to a brown-red liquid, from which crystals of *thiazyl chloride* (NSCl) form. It also unites with chlorides of S to crystalline compounds.

Imido-sulphonic acid, $\text{H}-\text{N} \begin{matrix} \text{S}-\text{O}-\text{O}-\text{O}-\text{H} \\ \text{S}-\text{O}-\text{O}-\text{O}-\text{H} \end{matrix}$, forms by action of NH_3 on SO_3 , or on sulphuryl oxychloride (SO_3HCl).

Amido-sulphonic acid ($\text{H}_2=\text{N}-\text{S}-\text{O}-\text{O}-\text{O}-\text{H}$).

XXXIV. SELENIUM.

Symbol, Se. Diad and hexad (Se^{II} and Se^{VI}). Atomic weight, 78.797 . Specific gravity, 4.28 (of crystals, 4.50).

Occurs with native S in pyrites (also as CuSe ; PbSe ; Ag_2Se , and HgSe). Made from deposits of some Pb chambers and from chimney dust of H_2SO_4 works.

Properties: Melts at 217°C ., boils at 676° to 683°C . with reddish vapor. When melted and quickly cooled it forms *amorphous* Se, which is reddish-black, translucent, brittle; specific gravity, 4.28 . Dissolves

in CS_2 , from which red crystals deposit; also in K_2Se , from which black crystals form. Dissolves in cold concentrated H_2SO_4 ; precipitated by H_2O . Burns in air with radish odor. Heated to 97°C . it suddenly becomes *crystalline*; dark-gray, metallic lustre; conducts electricity (amorphous does not) when exposed to light—the greater the light the greater its conducting power (application to telephone); insoluble in CS_2 .

Selenium hydride (H_2Se): Made from FeSe and HCl . Colorless poisonous gas, offensive odor; soluble in H_2O , decomposed in air. At 150°C . it begins to dissociate and continues up to 270°C ., when dissociation becomes feeble and ceases at 520°C .

Selenium dichloride (Se_2Cl_2): Brownish-yellow oil, decomposed by H_2O . Made like S_2Cl_2 .

Selenium tetrachloride (SeCl_4): White needles.

Selenyl chloride (SeOCl_2).

Selenium sulphide (SeS): Formed in volcano, Lipari Islands.

Selenious acid (H_2SeO_3) and *selenium dioxide* (SeO_2): H_2SeO_3 is formed by dissolving Se in concentrated HNO_3 , or by burning Se in O to SeO_2 and dissolving in H_2O . Large, colorless crystals.

Selenic acid (H_2SeO_4): Not directly obtainable by oxidation of H_2SeO_3 (no SeO_3 known) but by passing Cl through silver selenite (Ag_2SeO_3). Decomposed at 280°C .

Selenium nitride (SeN): Orange-yellow, explosive powder.

XXXV. TELLURIUM.

Symbol, Te . Diad and hexad (Te^{II} and Te^{VI}). Atomic weight, 127.96. Specific gravity, 6.398 at 20°C . Melts, 452° to 455°C .

Discovered by Müller von Reichenbach, 1782. First studied by Klaproth, 1798.

Occurs with Au , Si , Pb , Bi , and Hg (coloradoite) in Transylvania, Colorado, and Bolivia.

Properties: Silvery, metallic lustre; brittle, rhombohedric crystals. Dissolves in H_2SO_4 with purple color; precipitates unchanged by diluting with H_2O . Burns with blue flame and no odor in air to TeO_2 . Distills above melting point. Specific gravity of vapor at $1,300^\circ \text{C}$. is 8.743 (air = 1), or 126.0 ($\text{H} = 1$). *Amorphous* (specific gravity, 5.928) Te is a poor conductor of electricity.

Tellurium hydride (H_2Te): Colorless, poisonous, odorous gas; slightly soluble in H_2O ; easily decomposed. *Tellurium dichloride* (TeCl_2): Black crystals; melt at 209°C . *Tellurium tetrachloride* (TeCl_4): White crystals; melt at 224°C . and boil at 414°C . with yellow vapor. *Tellurium dibromide* (TeBr_2) is black, and the *tetrabromide* (TeBr_4) is orange.

Tellurium dioxide (TeO_2): Made by burning Te in O; octahedra; insoluble in H_2O ; distills without change.

Telluric acid (H_2TeO_3): Made by dissolving Te in HNO_3 ; separates on addition of H_2O ; also from TeCl_4 and H_2O ; decomposes easily into TeO_2 and H_2O .

Tellurium trioxide (TeO_3): Made by cautiously heating telluric acid; yellow, amorphous, insoluble in H_2O .

Telluric acid (H_2TeO_4): Easily reduced to Te by glucose; with H_2O forms crystals of $\text{H}_2\text{TeO}_4 + 2\text{H}_2\text{O}$.

Tellurium disulphide (TeS_2): Brown precipitate by H_2S from potassium tellurate.

XXXVI. NITROGEN GROUP.*

N; P; As (Sb; Bi; Vd; Id; Nb; Ta). See page 90.

Triatomic with H. Pentatomic and triatomic with O and Cl groups. Oxides are acid with first of series, becomes more basic with increase of atomic weight.

XXXVII. PHOSPHORUS.

Symbol, P. Triad and pentad (P^{III} and P^{V}). Atomic weight, 30.96 ($\text{P}_4 = 123.84$). Specific gravity, 1.83. Melts, 44°C . Boils, 290°C .

Discovered in urine by Brand, of Hamburg, 1669. Made from bones by Gahn and Scheele. Specific gravity of vapor, 61.92 ($\text{H} = 1$).

Occurs as phosphate in phosphorite; as apatite; in bones as Ca salts, etc.

Properties: Several modifications. 1. *Ordinary*, colorless or yellowish; transparent, waxy, or brittle when cold. One part P dissolves in 350 absolute alcohol, 240 boiling absolute alcohol, 80 absolute ether, 50 fatty oils, and 18 in 1 of CS_2 ; insoluble in H_2O . Volatile when distilled with H_2O ; colorless vapor. Rhombic crystals from solution in CS_2 . When kept under H_2O in light, crystallizes and becomes opaque. Ignites at 50°C . When burned with excess of O, forms P_2O_5 , or, when O is deficient, P_2O_3 . Ignites in Cl forming PCl_3 . Luminous by oxidation. Strong reductent; separates Au, Ag, and Pb from solutions. Poisonous. 2. *Red (amorphous)*, formed from ordinary by action of direct sunlight or heating to 250°C . in closed vessel; rectified by heating with NaOH. Scarlet-red powder, sometimes crystalline; specific gravity, 2.15 to 2.34; unchanged in air; not luminous; insoluble in CS_2 ; does not melt at 300°C .; not poisonous; ignites at 260°C .; more easily formed in presence of S, Se, or Te. 3. *Black*; formed by heating P in vacuum to higher temperature;

*Phosphorus group.

black rhombohedral crystals; specific gravity, 2.34; also formed in melted Pb. 4. *Liquid?* (E. J. Houston.) Match manufacturers are poisoned with P; dangerous wounds by burns; oil of turpentine and CuSO_4 are antidotes; caution in prescribing P; greenish coloration of flame; spectrum has two green lines.

Manufacture: Generally from bone ash containing about 80 per cent of $\text{Ca}_3(\text{PO}_4)_2$: $3[\text{Ca}_3(\text{PO}_4)_2] + 6\text{H}_2\text{SO}_4 = 3[\text{CaH}_4(\text{PO}_4)_2] + 6\text{CaSO}_4$; the P salt is extracted with H_2O , leaving the Ca salt behind; the solution is evaporated and mixed with powdered charcoal, dried and distilled in clay retorts: $3[\text{CaH}_4(\text{PO}_4)_2]$, heated, becomes $3[\text{Ca}(\text{PO}_3)_2] + 6\text{H}_2\text{O}$; with distilled with charcoal, $2[\text{Ca}(\text{PO}_3)_2] + 5\text{C} = \text{P}_2 + \text{Ca}_2\text{P}_2\text{O}_7 + 5\text{CO}$. Only one-half of the P is thus obtained in the free state.

Uses: In manufacture of matches; in medicine for nervous exhaustion, debility, etc. (dose, $\frac{1}{100}$ to $\frac{1}{75}$ grain).

Oleum phosphoratum (U. S. P.): P, 1 part; ether, 9 parts; oil of sweet almond, sufficient to make 100 parts; dose, 3 to 5 minims.

Pilule phosphori (U. S. P.): One grain in 100 pills; protected from oxidation by coating of tolu.

P unites directly with many elements producing heat and light; hence, it must be kept carefully under H_2O . When partially covered, P_2O_3 and O_3 and other compounds form. Light readily affects P.

Experiments: Burning P in O. Melting P in hot H_2O . Dissolving P in CS_2 and putting on paper; self-ignition. Heating red P. Safety matches.

Tests: Luminosity on distilling with H_2O (prevented by presence of turpentine, alcohol, ether, or butyric acid). Passing H through warm solution, burns with green flame.

Phosphorus with hydrogen: Gaseous phosphine (PH_3); liquid (P_2H_4 ; $\text{H}_2\text{P}-\text{PH}_2$); solid (P_4H_2 ; $\text{HP}=\text{P}-\text{P}=\text{PH}$); phosphonium (PH_4).

Gaseous phosphine (PH_3): P does not unite directly with H, but when P is heated with aqueous solution of potassa, PH_3 results: $3\text{P} + 4\text{KOH} + 2\text{H}_2\text{O} = \text{PH}_3 + \text{H}_2 + \text{K}_3\text{PO}_4 + \text{KH}_2\text{PO}_2$. Also prepared with P_2H_4 by Ca_3P_2 and H_2O ; pure by PH_4I and H_2O . A colorless gas with garlicky odor; insoluble in H_2O , soluble in alcohol and ether; very combustible. Not self-inflammable when pure, but becomes so when mixed with P_2H_4 (burning to P_2O_5 and H_2O with formation of rings) or when mixed with HNO_2 . Sunlight, S or C destroys self-inflammability. Very poisonous. Decomposed with explosion by Cl; $\text{PH}_3 + \text{Cl}_2 = \text{PCl}_3 + 3\text{HCl}$. Decomposed by silver nitrate. Unites with HI to form phosphonium iodide (PH_4I), in colorless, transparent, cube-like rhombohedra; sublimes unchanged, but decomposes by H_2O . PH_3 also forms compounds with HBr; SnCl_4 ; SbCl_5 , and TiCl_4 .

Liquid phosphine (P_2H_2): May be separated by cooling the self-inflammable gas; forms a colorless liquid; insoluble in H_2O ; inflames

in contact with O; gradually separates into PH_3 and P_4H_2 : $5\text{P}_2\text{H}_4 = 6\text{PH}_3 + \text{P}_4\text{H}_2$.

Solid phosphine (P_4H_2): Also formed by decomposing Ca_3P_2 by warm concentrated HCl. Yellow flocculent powder; ignites by heating in air to 160°C .

Chlorides of phosphorus: Trichloride (PCl_3), pentachloride (PCl_5), and oxychloride (POCl_3).

PCl_3 made by burning P in Cl. A colorless liquid; specific gravity, 1.616 at 0°C .; boils at 76°C .; fumes in moist air, forming H_3PO_3 and HCl; good solvent for P.

PCl_5 made by addition of dry Cl to PCl_3 or to solution of P in CS_2 . Colorless crystals; sublimes at 100°C .; forms PCl_3 and Cl at higher temperatures.

POCl_3 made with HCl from PCl_5 with a little H_2O . A colorless liquid; boils at 110°C .; decomposes with much H_2O .

Bromides of phosphorus:

Tribromide (PBr_3): colorless liquid; specific gravity, 2.928; boils at 175°C .

Pentabromide (PBr_5): yellow crystals, easily decomposed.

Oxybromide (POBr_3): colorless crystals; fuses at 45°C .; boils at 195°C . PCl_3Br_2 ; orange solid. PCl_3Br_4 ; dark-red crystals, red by reflected light. PCl_3Br_8 ; brown needles, green metallic lustre; melts at 25°C .; decomposes at 90°C . PCl_2Br_7 ; unstable prismatic crystals.

Iodides of phosphorus:

Tri-iodide (PI_3): Made by action of 12 parts I on 1 part P dissolved in CS_2 ; dark-red crystals; melts at 55°C .

Tetra-iodide (P_2I_4): Made by action of 41 parts I on 5 parts P dissolved in CS_2 ; orange prisms; melts at 110°C . PCl_3I_2 ; made by digestion of PCl_3 with I; red six-sided crystals.

Fluorides of phosphorus:

Trifluoride (PF_3): Made from AsFe and PCl_3 ; colorless liquid; boils at 60°C .

Pentafluoride (PF_5): Made from AsF₅ and PCl_5 ; colorless gas; fumes in air; specific gravity is 63.534 ($\text{H} = 1$); decomposed by H_2O .

Compounds with halogens more stable, the less their atomic weight.

Oxides of phosphorus:

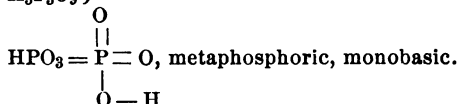
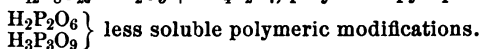
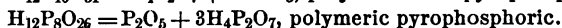
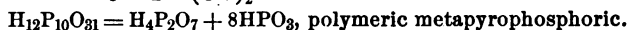
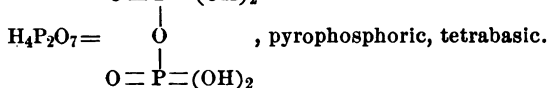
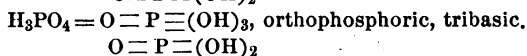
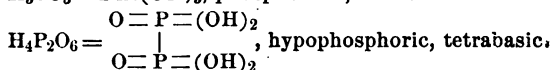
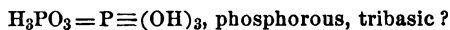
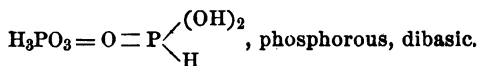
P_2O , monoxide.

P_2O_3 , trioxide.

P_2O_5 , pentoxide.

Acids of phosphorus:

$\text{H}_3\text{PO}_2 = \text{O} = \text{P} \begin{matrix} \text{OH} \\ \text{H}_2 \end{matrix}$, hypophosphorous, monobasic.



Hypophosphorous acid (H_3PO_2 ; $\text{H} \cdot \text{PH}_2\text{O}_2$): Boil P with $\text{Ba}(\text{OH})_2$ and add H_2SO_4 to set H_3PO_2 free. Salts very soluble. Acid forms white crystals; melts at $174^\circ \text{C}.$; very soluble in H_2O ; boiling decomposes it into PH_3 and H_3PO_4 ; $2\text{H}_3\text{PO}_2 = \text{PH}_3 + \text{H}_3\text{PO}_4$; strong reductent, reduces metals; also H_2SO_4 to SO_2 , and SO_2 to S. Salts absorb O and become phosphates. On heating forms PH_3 and metaphosphates or phosphides.

Phosphorous acid (H_3PO_3): There are most likely two acids; the dibasic ($\text{H}_2 \cdot \text{PHO}_3$) and tribasic (H_3PO_3) the latter as yet not studied. Formed by contact of P with moist air or by PCl_3 and H_2O . Colorless crystals; very soluble in H_2O ; absorbs O and becomes H_3PO_4 ; melts at $70^\circ \text{C}.$; decomposes at $180^\circ \text{C}.$ into PH_3 and H_3PO_4 ; $4\text{H}_3\text{PO}_3 = \text{PH}_3 + 3\text{H}_3\text{PO}_4$. Strong reductent.

Phosphorus trioxide (P_2O_3): Formed by imperfect oxidation of P in cool air or O, or by $\text{PCl}_3 + \text{H}_3\text{PO}_3$. White hygroscopic powder; attracts O from air, and burns to P_2O_5 and H_3PO_4 ; dissolves in H_2O to H_3PO_3 .

Hypophosphoric acid ($\text{H}_4\text{P}_2\text{O}_6$): Formed with others by slow oxidation of P. Na salt ($\text{Na}_2\text{H}_2\text{P}_2\text{O}_6$) little soluble. Insoluble Pb salt from which free acid is obtained by H_2S . Monoclinic prisms.

Orthophosphoric acid (H_3PO_4): Made (U. S. P. process) by gently heating P in HNO_3 , or P_2O_5 and HNO_3 and H_2O ; commercially from bone ash by H_2SO_4 . Contains As from impure P. Purified by H_2S . Colorless prisms; melts at $41.7^\circ \text{C}.$; inodorous; very sour; deliquescent; does not coagulate albumen. Yellow Ag salt. Forms

three series of salts: 1, *primary* (NaH_2PO_4); 2, *secondary* (Na_2HPO_4); 3, *tertiary* (Na_3PO_4). *Acidum phosphoricum* (U. S. P.) = 50 per cent H_3PO_4 + 50 per cent H_2O ; dose, 5 to 10 minims. *Acidum phosphoricum dilutum* (U. S. P.) contains 5 per cent H_3PO_4 .

Pyrophosphoric acid ($\text{H}_4\text{P}_2\text{O}_7$): Made by heating *secondary ortho-phosphate* or *ortho acid* to 200° to 300° C.; colorless crystals; soluble in H_2O ; does not coagulate albumen. White Ag salt: It is changed to *ortho acid* by long boiling with H_2O . It also forms *amido acids* [$\text{P}_2\text{O}_3(\text{OH})_2(\text{NH}_2)_2$ and $\text{P}_2\text{O}_3\cdot\text{OH}(\text{NH}_2)_3$].

Metaphosphoric acid (HPO_3): Made by dissolving P_2O_5 in cold H_2O ; $\text{P}_2\text{O}_5 + \text{H}_2\text{O} = 2\text{HPO}_3$; also by heating H_3PO_4 or $\text{H}_4\text{P}_2\text{O}_7$ above 400° C. *Glacial acid* of commerce is glassy, transparent, but not pure; melts and volatilizes undecomposed by heat; deliquescent; very soluble in H_2O ; coagulates albumen; boiling with H_2O forms H_3PO_4 . White Ag salt.

Phosphorous pentoxide (phosphoric acid anhydride, P_2O_5): White, amorphous solid, formed by combustion of P in O or air; sublimes unchanged at high temperature; very hygroscopic; hisses with H_2O ; used for drying gases, etc.

Chloride of pyrophosphoric acid ($\text{P}_2\text{O}_5\text{Cl}_4$): Colorless liquid; fumes in air; specific gravity is 1.58 at 7° C.; boils at 210° to 215° C.; forms HCl and H_3PO_4 with H_2O .

Compounds of phosphorus and sulphur: Trisulphide (P_4S_3); hexasulphide (P_8S_6); pentasulphide (P_2S_5).

S and P, when mixed under H_2O , liquefy and form mixture from which excess of S or P crystallizes. Explodes violently on heating. Best made with red P heated in CO_2 with calculated weight of S.

P_4S_3 is soluble in CS_2 ; yellow rhombic prisms; melts at 166° C.; boils at 100° C.

P_8S_6 melts at 296° to 298° C.; yellow transparent needles.

P_2S_5 melts at 274° to 276° C., boils at 530° C.; pale-yellow crystals; specific gravity is 7.6; decomposes with H_2O .

PSCl_3 , PSBr_3 , and P_2SBr_4 also exist.

Phospham (PN_2H_2): Made from NH_3 and PCl_3 ; white powder, insoluble in H_2O .

XXXVIII. ARSENIC.

Symbol, As. Triad and pentad (As^{III} and As^{V}). Atomic weight, 74.92 ($\text{As}_4 = 299.68$). Specific gravity, 3.69 to 3.74. Sublimes, 200° C.

Occurs native as cobaltum or as sulphide (realgar, auripigmentum), or with metals (speisscobalt, Cu, Ni, mispickel, etc.). Obtained by heating arsenical pyrites (mispickel) or arsenious acid with charcoal.

Properties: Two modifications similar to P. *Amorphous*, black,

glassy mass; specific gravity 4.71; very brittle. *Crystalline* in acute hexagonal rhombohedra; strong metallic lustre; specific gravity 5.727. At ordinary temperatures As volatilizes without melting, but melts under pressure. Vapor density at $860^{\circ}\text{C.} = 149.84$ (*four atoms in a molecule*); density less at high heat. When sublimed in H, crystals form near heated part of tube; amorphous in cooler part, and whenever vapor is cooled to 210° to 220°C. At 360°C. amorphous changes to crystalline. Vapor is lemon yellow. Oxidizes in moist air (amorphous less easily than crystalline). Burns in O with a garlicky odor to As_2O_3 . Unites directly with most elements. Not affected by HCl; oxidized by HNO_3 . In some metallic compounds it replaces S in atomic ratio (FeS_2 to FeAs_2).

Uses: Shot rendered spherical by addition of As. Fly poison. Pyrotechnics. Poisonous. Insoluble in H_2O until oxidized. Oxides used for many purposes, anilin, glass-making, etc.

Cobaltum, a misnomer.

Arsine (AsH_3): Formed from all compounds of As by H. Colorless, odorous, poisonous gas; specific gravity, 77.92 ($\text{H} = 1$); liquefies at -40°C. to colorless liquid; no basic properties; little soluble in H_2O . Burns with bluish-white flame to As_2O_3 and H_2O , unless a cold body held in the flame prevents the burning of the As, which then deposits as a mirror while only H burns. Decomposed by red-heat into As and H; also by oxidizing bodies (Cl, Br, I, HNO_3 , etc). Reduces AgNO_3 to Ag (forming As_2O_3) and CuSO_4 to Cu_2As_3 . *Marsh's test* is based upon this behavior of AsH_3 ; the As spots are brownish and lustrous, and disappear with NaOCl , while Sb spots are black, velvety, and do not disappear with NaOCl . AsH_3 also made by adding potassa and Al to As solutions. No analogue to P_2H_4 is known, but its methyl derivative is *kakodyl* $(\text{CH}_3)_2 = \text{As}-\text{As} = (\text{CH}_3)_2$.

Solid arsine (As_4H_2): Reddish powder, decomposed by heat, formed by nascent H in presence of HNO_3 .

Arsenic trichloride (AsCl_3): Made by burning As in Cl, or distilling As_2O_3 with concentrated HCl. Colorless, fuming liquid; boils at 130.2°C. ; specific gravity is 2.205; very poisonous; mixes with little, but decomposes with much H_2O to HCl and As_2O_3 ; dissolves P and S unchanged.

Arsenic hydroxylchloride [$\text{As}(\text{OH})_2\text{Cl}$]: Crystals deposited from strong aqueous solution of AsCl_3 .

Arsenic oxychloride (AsOCl): Formed by mixing boiling AsCl_3 with As_2O_3 . A tough, amorphous mass.

Arsenic tribromide (AsBr_3): Colorless prisms; melts at 20° to 25°C. ; boils at 220°C. ; decomposes by much H_2O .

Arsenic tri-iodide (AsI_3): Brick-red hexagonal plates; melts at 146°C. ; boils at 394° to 414°C. ; made by heating As in solution of I in

ether, or by subliming mixture of As and I; soluble in 3.5 parts H_2O , in 10 parts alcohol, in ether and in CS_2 .

Arsenic trifluoride (AsF_3): Made by distilling 1 part CaF_2 and 1 part As_2O_3 with 5 parts concentrated H_2SO_4 ; colorless, fuming liquid; specific gravity is 2.66 at 0°C .; boils at 60.4°C .; decomposed by H_2O .

Arsenic pentafluoride (AsF_5): Only known in compounds; arsenofluoride of potassium $(\text{KF} \cdot \text{AsF}_5)_2 + \text{H}_2\text{O}$.

Arsenic oxyfluorides: AsOF_3 and As_2OF_6 in compounds with KF .

Compounds of arsenic and oxygen: As_2O_3 ; As_2O_5 ; H_3AsO_3 ; HAsO_2 ; H_3AsO_4 ; $\text{H}_4\text{As}_2\text{O}_7$; HAsO_3 .

Arsenic trioxide (As_2O_3): Occurs native as Arsenic bloom. Formed by oxidation of As or of pyrites. On large scale it is collected in chambers and rectified in iron retorts. Commercial is *amorphous* and glassy; specific gravity, 3.69; but gradually changes to *crystalline*, opaque, porcelain-like; specific gravity, 3.74. Crystals heated near 200°C . for some time become glassy. Vapor density from 600° to $1,500^\circ \text{C}$. = 198.9, corresponding to As_4O_6 . The crystals are *dimorphous*; regular octahedra from condensed vapor or solution in dilute HCl ; rhombic prisms from roasting furnaces and alkaline solutions. Boiling converts crystalline into amorphous, which is more soluble than the former. Dissolves slowly in cold liquids, so that solid As_2O_3 is long preserved. In twenty-four hours there dissolves 1 part amorphous in 108 of H_2O , or 1 part crystalline in 355 of H_2O , forming acid solution. One part amorphous will dissolve in 30 of H_2O , or 1 of crystalline in 80 of H_2O at 15°C . One part of either dissolves in 15 of boiling H_2O . Sparingly soluble in alcohol, freely in HCl , moderately in glycerin. *Poisonous* in small doses. Medicinal dose about $1/20$ grain three times a day; fatal dose, from $1/2$ to 2 grains. Tolerance: As eating in Styria, etc. *Symptoms of poisoning*: Metallic taste, ptyalism, constriction of pharynx, burning in stomach, nausea, retching, vomiting of brown or bloody matter, black and fetid stools, thirst, etc., collapse. *Post-mortem signs*: Inflamed mucosa; red, velvety rugæ of stomach; eschars. *Elimination* by urine. *Antidotes*: Ferric hydrate, magnesium hydrate.

Arsenious and metarsenious acids (H_3AsO_3 and HAsO_2): Not known in pure state but in salts; yellow Ag_3AsO_3 , green CuHAsO_3 , white PbHAsO_3 , white $\text{Pb}(\text{AsO}_2)_2$. Salts called *arsenites*.

Arsenic pentoxide (As_2O_5): Made by carefully heating H_3AsO_4 to low red-heat. Colorless, glassy solid; dissolves in H_2O , forming H_3AsO_4 . High heat decomposes it into As_2O_3 and O_2 . With HCl it forms AsCl_3 and Cl , and H_2O .

Orthoarsenic acid (H_3AsO_4): Occurs in minerals (Ca, Pb, Ni). Made by passing Cl through H_2O containing As_2O_3 in suspension or by oxidizing As_2O_3 by hot concentrated HNO_3 or aqua regia. White

crystalline mass, rhombic deliquescent crystals $(\text{H}_3\text{AsO}_4)_2 + \text{H}_2\text{O}$, loose H_2O at 100°C . Reduced by SO_2 to As_2O_3 , and by nascent H to AsH_3 , and by fuming HCl to $\text{AsCl}_3 + \text{Cl}_2$. Strong tribasic acid, similar to H_3PO_4 ; salts are isomorphous and called *arsenates*. Reddish-brown Ag salt, white NH_4 and Mg salts.

Pyroarsenic acid ($\text{H}_4\text{As}_2\text{O}_7$): *Made* by evaporating solution of H_3AsO_4 at 180°C ., and letting concentrated residue cool. Hard, shining, prismatic crystals.

Metarsenic acid (HAsO_3): *Made* by heating $\text{H}_4\text{As}_2\text{O}_7$ for some time at 200°C ., and then slowly raising to 206°C . Addition of H_2O restores $\text{H}_3\text{As}_3\text{O}_4$.

Arsenic trisulphide (As_2S_3): Natural and artificial. Precipitated from acid solutions of arsenates by H_2S . Soluble in alkaline sulphides, NH_4OH and $(\text{NH}_4)_2\text{CO}_3$. Rhombic prisms. Melts to red mass. Forms *sulpharsenites*.

Arsenic pentasulphide (As_2S_5): *Made* by precipitating arsenates by H_2S from acid solutions. Yellow powder. Forms *sulpharsenates* from which HCl precipitates H_3AsS_4 .

Arsenic with phosphorus: *Made* by passing AsH_3 through PCl_3 , or PH_3 through AsCl_3 . Red-brown powder, decomposed by H_2O . Burns in air.

Tests for arsenic:

1. Acid solution of $\text{As} + \text{H}_2\text{S} =$ yellow As_2S_3 ; soluble in $(\text{NH}_4)_2\text{S}$, NH_4OH , or $(\text{NH}_4)_2\text{CO}_3$; CdS is not. Detects 1 part As in 8,000.

2. *Reinsch's test*: Spot of As on Cu foil by heating As solution with HCl ; liability to mistake Hg or Ag . Detects 1 part As in 40,000.

3. *Hume's test*: Slightly ammoniacal solution of $\text{As} +$ silver nitrate = yellow Ag_3AsO_3 .

4. *Scheele's test*: Slightly ammoniacal solution of $\text{As} +$ cupric sulphate = green CuHAsO_3 .

5. *Marsh's test*: AsH_3 burned forms As mirror. Liability to mistake Sb . Detects 1 part As in 200,000,000.

6. *Reduction test*: Reduction of As compound with charcoal forms As mirror on glass. Sublimation forms octahedra.

7. Melting with K or Na acetate forms *kakodyl*.

8. *Odor test*: Garlicky odor on reduction.

XXXIX. SILICON.

Symbol, Si . Tetrad (Si^{IV}). Atomic weight, 28.2. Specific gravity, 2.49. Never occurs uncombined, but next to O is the most abundant element. *Made* by heating Na with SiF_4 or SiCl_4 , or by heating Na_2SiF_6 with Na and Zn ; amorphous by first and crystalline by second process. Si is analogous to C , also to Sn , Ti , Zn , and Th . Amor-

phous Si is a brown powder; burns in O to SiO_2 . *Crystalline Si* in regular, brittle, hard octahedra; may be heated in O without change; melts at high temperatures; not affected by acids; burns in Cl; dissolves in caustic alkalies. A *graphitoid* variety is said to be formed by melting 10 parts cryolite and 5 of potassium silicate with 1 of Al.

Silicon and hydrogen compound (SiH_4): Made by dissolving Mg_2Si in HCl. Colorless gas; liquid at -50°C . under 50 atmospheres, or at -1°C . under 100 atmospheres pressure; self-inflammable when heated or containing free H; explodes with Cl.

Silicon and halogen compounds: SiCl_4 ; SiBr_4 ; SiI_4 ; SiF_4 ; SiCl_3Br ; SiBr_3I ; H_2SiF_6 ; Si_2Cl_6 ; Si_2Br_6 ; Si_2F_6 ; Si_2I_6 ; SiHCl_3 ; SiHBr_3 ; SiHI_3 ; Si_2OCl_6 .

Silicon tetrachloride (SiCl_4): Made by heating Si, or finely powdered SiO_2 and charcoal, in Cl. Colorless, mobile liquid; specific gravity = 1.52; boils at 57.6°C .; vapor density, 84.84; fumes in air; decomposed by H_2O to HCl and SiO_2 .

Silicon hexachloride (Si_2Cl_6): Colorless liquid, forming crystals below 0° . Constant under 350° and over $1,000^\circ \text{C}$., but dissociated between these temperatures.

Silicon tetrafluoride (SiF_4): Made by heating equal parts of powdered fluor spar (or cryolite) and quartz with an excess of H_2SO_4 . Colorless, pungent gas; condensed at low temperatures; decomposed by H_2O into SiO_2 and H_2SiF_6 ; $3\text{SiF}_4 + 2\text{H}_2\text{O} = \text{SiO}_2 + 2\text{H}_2\text{SiF}_6$.

Hydrofluosilicic acid (H_2SiF_6): Dibasic; precipitates Ba and K salts; easily decomposed at high temperatures.

Silicon chloroform, SiHCl_3 , also *bromoform* and *iodoform*.

Silicon and oxygen compounds:

Silicon dioxide (SiO_2): Made from any silicic acid by sufficient heat. Occurs in nature as quartz, amethyst, tridymite, etc. Forms hexagonal prisms and pyramids. Hardness = 7. Specific gravity = 2.3 to 2.6. Melts in compound blow pipe to transparent glass. *Amorphous* SiO_2 is soluble in alkalies.

Silicic acids (H_2SiO_3 ; H_2SiO_4 and polyacids): H_2SiO_3 is a gelatinous mass, from decomposition of alkaline silicates. H_2SiO_4 made by dialysis from H_2SiO_3 . Heat converts H_2SiO_4 into polyacids; $\text{Si}_2\text{O}_3(\text{OH})_2$; $\text{Si}_3\text{O}_4(\text{OH})_4$; $\text{Si}_3\text{O}_5(\text{OH})_2$; $\text{Si}_4\text{O}_7(\text{OH})_2$; etc. They form the natural silicates; opal, hyalith, etc.

Silicon formic acid anhydride, $\text{O} \begin{array}{l} \diagup \text{Si} - \text{O} - \text{H} \\ \diagdown \text{Si} - \text{O} - \text{H} \end{array}$.

Silicon oxalic acid, $\begin{array}{c} \text{Si} - \text{O} - \text{O} - \text{H} \\ | \\ \text{Si} - \text{O} - \text{O} - \text{H} \end{array}$.

Silicon bisulphide (SiS_2): Made by heating amorphous Si in S vapor. White, shining needles; sublimes; decomposed by H_2O .

Silicon chlor-hydro-sulphide (Silicon mercaptan, $\text{SiCl}_3\text{—SH}$): Colorless liquid; boils at 96° .

Silicon nitride (Si_2N_3)₂: Forms when N is passed over heated crystalline Si; white, infusible powder.

$\text{Si}_2\text{N}_3\text{H}$: A white mass; soluble in HF and alkalis.

XL. BORON.

Symbol, B. Triad (B^{III}). Atomic weight, 10.94; specific gravity, 2.68.

Found only in compounds (boric acid, borax, boracite, datolith, etc.).

Amorphous made by reducing B_2O_3 by Na. Greenish-brown powder; ignites when heated in air, forming B_2O_3 and BN.

Crystalline made by alloying with Al, and dissolving out the Al by HCl. Lustrous-brown to black quadratic crystals; hard as diamond, especially when Al or C is present as impurities; does not burn in air; unchanged in air, or by acids or alkalis.

Boron with hydrogen (BH_3 ?) : Unknown in pure state; made by mixing H with Mg boron and HCl_3 , or BCl_3 and hot Mg; gas burns with green flame; deposits B on heating.

Boron with halogens: BCl_3 ; BBr_3 ; BF_3 ; HBF_4 .

Boron trichloride (BCl_3) : *Made* by heating B, or B_2O_3 and C, in Cl. Mobile, colorless liquid; specific gravity = 1.35 at 17°C .; boils at 18°C .; vapor density = 58.52; fumes in air; decomposed by H_2O ; unites with NH_3 to crystals of $2\text{BCl}_3 \cdot 3\text{NH}_3$; unites with NOCl to colorless, rhombic octahedra of $\text{BCl}_3 \cdot \text{NOCl}$.

Boron tribromide (BBr_3) : Colorless liquid; boils at 90°C .; specific gravity = 2.69.

Boron trifluoride (BF_3) : *Made* by heating fluorspar and B_2O_3 to white heat or distilling 2 parts borax (or 1 of boric acid), 2 of fluorspar, and 12 of concentrated H_2SO_4 . Colorless gas; specific gravity = 38.94; liquid at low temperatures.

Hydrofluoboric acid (HBF_4) : *Made* when BF_3 is absorbed by H_2O . Forms *borofluorides*.

Boron with oxygen: B_2O_3 ; H_3BO_3 ; HBO_2 .

Boric acid (H_3BO_3) : *Occurs* in volcanic regions in Tuscany; either solid, as *sassolin*, or in *suffioni* (volcanic vapors and lagoons). Colorless, shining, triclinic scales; soluble in 25 parts of H_2O at 15°C . or 3 at 100°C .; soluble in alcohol; acid reaction; volatile with steam; colors flame green; antiseptic; is converted into $\text{H}_2\text{B}_4\text{O}_7$ at 100°C ., and $\text{H}_2\text{B}_6\text{O}_{10}$ at 140°C ., and finally into B_2O_3 ; expels even H_2SO_4 at red heat.

Boron trioxide (B_2O_3) : *Made* by heating and melting H_3BO_3 . Color-

less, transparent glass; specific gravity = 1.8; soluble in H_2O and alcohol.

Boron sulphide (B_2S_3): *Made* by heating amorphous B in S vapor, or B_2O_3 and C in CS_2 . White, glassy mass; penetrating odor; decomposed by H_2O .

Boron nitride (BN): *Made* by heating amorphous B in N, or H_3BO_3 with urea, etc. White, amorphous, gray mass; unaffected by ignition; highly luminous in flame; insoluble in H_2O ; when heated with H_2O to $200^\circ C$. forms H_3BO_3 and NH_3 .

Boron and Cu form a bronze of similar properties to steel.

XLI. METALS.

No sharp line of demarkation exists between metals and metalloids. Some members of each division stand near the border, and have transitional characters.

Characteristics of metalloids:

They usually unite with H (also H and O). and form acids. With O they form *acid anhydrides* (acid oxides). An exception is H_2O (H acts as a metal). B forms no H compounds.

Characteristics of metals:

They form few H compounds (Cu_2H_2 ; Pd_2H ; K_2H). The O and OH compounds are mostly basic. An increase of O diminishes the basic properties.

The surface gives *metallic* lustre by reflected light. Except in thin laminæ they are opaque (gold leaf an example).

The color is a peculiar gray of various tints from white to blue; also pink (Bi), red (Cu), and yellow (Au).

The crystals are mostly in the regular system (cubes and octahedra). Those of more metalloid character form exceptions; Sn is quadratic (amalgam of Sn and Hg is regular), Sb and Bi are hexagonal rhombohedra with angles of near 90° , so as to resemble cubes, *As = $85^\circ 4'$, Te = $86^\circ 57'$, Sb = $87^\circ 35'$, and Bi = $87^\circ 40'$.

The specific gravity varies greatly. Those under 5 are called *light* metals.

Iridium.....	†21.800	Arsenic.....	*5.88 to 5.900
Platinum.....	21.1 to 21.500	Gallium.....	5.900
Uranium.....	18.400	Aluminium.....	2.56 to 2.670
Mercury.....	13.596	Magnesium.....	1.750
Lead.....	11.3 to 11.450	Rubidium.....	1.520
Bismuth.....	9.8 to 9.900	Potassium.....	0.865
Nickel.....	8.800	Osmium.....	†21.400
Cobalt.....	8.540	Gold.....	19.3 to 19.050
Manganese.....	8.000	Wolframium.....	17.600
Indium.....	7.420	Palladium.....	11.3 to 11.800
Zinc.....	6.86 to 7.100	Silver.....	10.500

* A metalloid.

† This heavy only in one modification.

Copper	8·960	Chromium	5·900
Cadmium	8·700	Barium	4·300
Molybdenum	8·630	Strontium	2·540
Iron	7·790	Calcium	1·580
Tin	7·300	Sodium	0·974
Antimony	6·7 to 6·800	Lithium	0·593

Also see table on page 10.

The hardness varies from extreme softness to great hardness. See scale of hardness on page 20.

Some metals are brittle and easily broken by compressing, pulling, or bending (Sb).

Some metals are ductile, and yield to pull, combined with lateral compression, forming a wire: Fe = 26, Cu = 17, Pt = 13, Ag = 8·9, Zn = 8, Au = 5·6, Sn = , Pb = 1.

Metals fuse below 100° C., as follows:

	Centigrade.		Centigrade.
Mercury	-39·0	Potassium	62·5
Gallium	30·1	Sodium	97·6
Rubidium	38·5		

Those fusing above 100° C. are:

	Centigrade.		Centigrade.
Lithium	180·0	Silver	957·0
Tin	227·8	Gold	1040·0
Cadmium	228·0	Copper	1150·0
Bismuth	258·0	Cast iron	1050·0 to 1200·0
Thallium	294·0	Steel	1300·0 to 1400·0
Lead	325·0	Wrought iron	1500·0 to 1600·0
Zinc	412·0	Palladium	1500·0
Antimony	425·0	Platinum	1800·0

Also see table on page 62.

Rate of conduction:

	Heat.	Electricity.
Silver	1,000	1,000
Copper	736	920
Gold	532	650
Iron	119	120
Zinc	...	240
Tin	145	140
Lead	85	80
Platinum	84	80
Bismuth	18	..
Mercury	..	2

The *magnetic* metals are: Fe; Ni; Co; Mn; Cr; Ce; Ti; Pd; Pe; and Os.

The *diamagnetic* metals are: Bi; Sb; Ag; Cu; Zn; Au; Sn; Cd; Hg; Na; Pb; As; Nr; Ir; and W.

The metals will not dissolve without forming chemical compounds. Many metalloids dissolve in water, alcohol, ether, carbon disulphide, etc., and may be recovered unchanged (S; Se; P; I; Br; Cl; O; H; N; etc.). A few metals dissolve in Hg, forming *amalgams*.

Some alloys (mixtures of metals) are real combinations, others only simple mixtures. Some crystallize and form compounds of greater hardness and lower fusing points than their constituents:

Bi.	Pb.	Sn.	Centigrade.
8 parts,	5 parts,	3 parts,	melts at.....100°0
2 parts,	1 part,	1 part,	melts at.....94°0
5 parts,	3 parts,	2 parts,	melts at.....93°6
4 parts,	2 parts,	1 part,	(Cd, 1 part), melts at.....70°0
.....1 part,	2 parts,	melts at.....182°0	
.....2 parts,	1 part,	melts at.....260°0	

Alloy of K and Na melts at 8° C.

Some metals *occur* native uncombined: Au; Ir; Os; Hg; Bi; Sb; Ag; Pt; Fe; Cu. Some occur as *ores* in combination with S, O, As, Si, etc.

Some metals absorb gases: Au; Ag; Pd; Fe.

Classification of Metals.

Groups by general resemblance:

Alkali metals.

Alkaline earth metals. } Lighter metals.

Earth metals.

Oxidizable, or ignoble metals.

Difficultly oxidizable, or noble metals. } Heavier metals.

Groups by isomorphism (forming compounds of equal crystalline forms):

K; Na; Rb; Cs; Li.

Ca; Mg; Mn; Fe^{II}; In; Co; Ni; etc.

Al^{III}; Fe₂^{III}; Cr₂^{III}; etc.

Groups by analytical reactions (behavior with H₂S and general reagents):

Alkalies.

Alkaline earths.

Earths.

Iron group.

Copper group.

Arsenic group.

Groups by atomicities:

Monads—K; Na; Li; Rb; Cs; Ag.

Diads—Ba; Sr; Ca; Mg; Cu; Zn; etc.

Triads—Au; Tl; Ga; In; Al (?).

Tetrads—Pt; Sn; Ti; Fe; Al; Pb.

Pentads—Bi; Sb*; As*; V; Ta; Nb.

Hexads—Cr; Nr; W; Mo.

Compounds of metals: Besides forming alloys and amalgams among themselves, metals unite with metalloids; Cu₂H₂ and Pd₂H.

Halogen compounds: Salts form according to the atomicities of the

*Metalloids.

metals. They are more constant than the halogen compounds of the metalloids, which are easily decomposed.

NaF; NaCl; NaBr; NaI.

CaF₂; CaCl₂; CaBr₂; CaI₂.

—F₃; AuCl₃; AuBr₃; AuI₃.

—F₄; PtCl₄; PtBr₄; PtI₄.

NbF₅; NbCl₅; —Br₅; —I₅.

—; —; TaBr₅; TaI₅.

Oxygen compounds: With O alone oxides form; with hydroxyl (OH) hydrates. The latter are formed like oxides, according to the atomicities of the metals; ($\equiv O^{II}$) forms K₂O; CaO; Fe₂^{III}O₃, etc. (—OH)^I forms KOH; Ca(OH)₂; Fe₂^{II}(OH)₆, etc. Some intermediate hydrates

form as $\begin{array}{c} \text{OH} \\ \diagup \quad \diagdown \\ \text{Cu} \quad \text{O} \\ \diagdown \quad \diagup \\ \text{Cu} \end{array}$ and $\begin{array}{c} \text{OH} \\ \diagup \quad \diagdown \\ \text{Au} \quad \text{O} \\ \diagdown \quad \diagup \\ \text{O} \end{array}$. Some hydrates are converted into

oxides by loss of OH₂, as Ca(OH)₂ = CaO + H₂O; Cu(OH)₂ = CuO + H₂O; Fe₂^{III}(OH)₆ = Fe₂^{III}O₃ + 3H₂O.

Salts with oxygen acids: These usually form by the metal replacing H atoms of the acid with liberation of H; K + HNO₃ = H + KNO₃; Zn + H₂SO₄ = H₂ + ZnSO₄, the acid radicle and the metal being united by the O of the OH; K + HO.NO₂ = H + KO.NO₂. Another mode of formation does not liberate H, but decomposes another molecule of acid, forming H₂O and an acid product of less oxygen; Cu + 2H₂SO₄ = CuSO₄ + 2H₂O + SO₂; Ag + 2HNO₃ = AgNO₃ + H₂O + NO₂. Some hydrates are converted into oxides by loss of OH₂, as in the following examples: Ca(OH)₂ = CaO + H₂O; Cu(OH)₂ = CuO + H₂O; Fe₂^{III}(OH)₆ = Fe₂^{III}O₃ + 3H₂O.

When oxides or hydrates form salts, the O or OH forms water with the H of the acid; KOH + HNO₃ = KNO₃ + H₂O.

Precedence of Bases.

Some acids have precedence over others, so as to displace them from their compounds in certain order; H₂SO₄ displaces HNO₃; HF, HCl, HBr, and HI take precedence in union with bases, so that HI is expelled by HBr; HBr by HCl, and HCl by HF.

Bases also precede each other. The first named metal in the following table displaces the second and subsequent ones, this displaces the third, etc., and so on through the list. Those in the first column liberate H from HCl and H₂SO₄. Those in the second column liberate SO₂ from H₂SO₄, in which they dissolve only when boiling acid is used. Those in the third column do not dissolve in even boiling H₂SO₄, but only in nitro-muriatic acid.

<i>H</i> liberated from H_2SO_4 .	SO_2 from boiling H_2SO_4 .	Not soluble in H_2SO_4 .
1. Potassium.	11. Lead.	17. Platinum.
2. Sodium.	12. Tin.	18. Gold.
3. Barium.	13. Bismuth.	
4. Strontium.	14. Copper.	
5. Calcium.	15. Mercury.	
6. Magnesium.	16. Silver.	
7. Aluminium.		
8. Zinc.		
9. Iron.		
10. Cadmium.		

When several acids as well as bases are concerned, the problem becomes more complicated. Thus : Potassium hydrate precipitates all of the succeeding metals from their chlorides or sulphates. Zinc precipitates those succeeding it in number; as Pb, Sn, Cu, etc.

Just as the same metals precede the others in electropositiveness, so every metal separates the more negative ones from its salts.

XLII. METALS OF THE ALKALIES.

	Li.	Na.	K.	Rb.	Cs.	(NH ₄) ?
Atomic weight.....	7.007	22.998	39.019	85.251	132.583	18.021
Specific gravity.....	0.5936	0.974	0.865	1.52	1.83	
Atomic volume.....	11.9	23.7	45.4	56.1	70.5	
Melting point; °C.	180.0	95.6	62.5	38.5	27.0	
Discovered.....	1817	1807	1807	1861	1860	
By.....	Arfved- son	Davy	Davy	B. & K.	B. & K.	
Calories liberated in forming chlorides.	93.700	97.700	105.600			

Characteristics.

All are monads.

All decompose cold water.

All have great affinity for O.

Hydrates and carbonates give alkaline reaction, and have caustic taste.

Salts are colorless unless formed by colored acids. All have low specific gravities. Oxides, hydrates, chlorides, bromides, iodides, fluorides, cyanides, sulphates, sulphides, carbonates, phosphates, and chlorates are all soluble.

XLIII. POTASSIUM.

Symbol, K. Monad. Atomic weight, 39.019. Specific gravity, 0.865. Melts, 62.5° C. Synonym, *kalium*.

Occurs only in combination; mostly as silicates with Al, etc.; as *feldspar* and other granitic compounds, from whose decomposition the

soil gets its K salts. Also occurs as chloride of potassium, in sea water and in mines. Found in plants and animals.

Prepared by electrolytic decomposition of caustic potash, or by white-hot iron and caustic potash. On large scale, by distilling a mixture of carbonate of potassium and C made by heating cream of tartar ($\text{KHC}_4\text{H}_4\text{O}_6$), and receiving the condensed metal under naphtha. *Purified* by pressing through cloth, under naphtha at 70°C ., or by re-distillation. Accidents may occur from the formation of explosive compounds with CO.

Properties of pure potassium: Silvery-white. Crystallizes in octahedra. Soft at ordinary temperatures. Specific gravity, 0.865. Melts at 62.5°C . Boils at 719° to 731°C . Vapor is green. Oxidizes in air. Burns with violet flame. Withdraws O from H_2O , even from ice, and ignites the escaping H.

Experiments with potassium: Lighting candle with icicle. Burning on water. Burning with chlorine. Explosion with bromine. Alloy with sodium in atomic proportions is liquid above 8°C . Amalgam crystallizes.

Potassium hydride (K_2H): When K is heated in H between 200° and 400°C . (best at 300°) it absorbs the gas and forms silvery-white, brittle, shining crystals, which remain unchanged in H atmosphere, but dissociate in vacuum.

Halogen Compounds of Potassium.

Potassium chloride (KCl): Occurs native as *sylvin* (near Stassfurt) in colorless cubes. Melts at 734°C ., and is volatile at higher temperatures. Soluble in 3 parts of H_2O at 15°C ., and in less of boiling H_2O . Insoluble in alcohol. Forms double salts with other chlorides: $\text{PtCl}_4 + 2\text{KCl}$; $\text{KCl} \cdot \text{ICl}_3$; etc.

Potassium iodide (KI): Made by neutralizing carbonate of potassium (K_2CO_3) with HI; by adding potassium carbonate (K_2CO_3) to ferrous ferric iodide (Fe_3I_8) made from 3 parts I, 1 part Fe filings in water and evaporating; or, by adding I to KOH, and, after evaporating solution, heating with C. Forms colorless or white porcelain-like cubes. Specific gravity is 2.9. Melts at 634°C ., and is volatile at higher temperatures. Water at 15°C . dissolves 1.3 parts, or at 100°C ., 2 parts. Soluble in 40 parts cold or 6 parts boiling alcohol (even in absolute alcohol). Its aqueous solution dissolves a large amount of I. From saturated solution prisms of $\text{KI} \cdot \text{I}_2$ form. Very deliquescent. Forms many double salts. When KI contains sodium iodide (NaI) it decomposes by CO_2 of air, forming sodium carbonate (Na_2CO_3) and HI, which turns the salt brown. KI has a neutral reaction.

Potassium bromide (KBr): Made by dissolving Br in KOH, and, after evaporation, heating with C; $6\text{KOH} + 3\text{Br}_2 = 5\text{KBr} + \text{KBrO}_3 + 3\text{H}_2\text{O}$. Forms colorless to nearly opaque white cubes. The specific gravity is 2.4. It melts at 699°C . Soluble in 1.6 parts water at 15°C .; 200 parts alcohol at 15°C .; 1 part boiling water; 16 parts boiling alcohol.

Potassium fluoride (KF): Forms cubes of $\text{KF} + 2\text{H}_2\text{O}$ at 15°C ., but anhydrous above 35°C . Melts at 789°C . Attacks glass. Forms double fluorides: KF.HF ; $2(\text{KF.AsF}_6) + \text{H}_2\text{O}$; 2KF.SbF_3 ; KF.BF_3 ; also the silicofluoride, 2KF.SiF_4 , which requires 833 parts of H_2O at 17°C . for solution.

Cyanogen Compounds of Potassium.

Potassium cyanide (KCN): Made by heating potassium ferrocyanide, $\text{K}_4\text{Fe}^{\text{II}}(\text{CN})_6$, to fusion, either alone, or the less pure with potassium carbonate (K_2CO_3).

Forms a white, crystalline mass of cubes. Is deliquescent, and very soluble in H_2O .

Used in photography, galvano-plastics, etc. At high temperatures it is a powerful reductent of metallic oxides. Is very poisonous.

Compound cyanides formed by KCN with metals:

$\text{K}_4\text{Fe}^{\text{II}}\text{C}_6\text{N}_6$, ferrocyanide.

$\text{K}_6\text{Fe}^{\text{III}}\text{C}_{12}\text{N}_{12}$, ferricyanide.

KCNS, sulphocyanide.

$\text{K}_2\text{Fe}(\text{NO})\text{C}_6\text{N}_5$, nitroprusside.

Oxides and Hydroxide of Potassium.

Potassium monoxide (K_2O): Made by heating K with KOH in H atmosphere. Forms a white, deliquescent, caustic mass. Volatile at red heat.

Potassium dioxide (K_2O_2): ?

Potassium tetroxide (K_2O_4): A yellow mass; made by heating K in dry O.

Potassium tetratoxide (suboxide, K_4O): ?

Potassium hydroxide (hydrate, caustic potash, KOH): Made by gradually adding $\frac{1}{2}$ part caustic lime, $\text{Ca}(\text{OH})_2$, to a boiling solution of 1 part potassium carbonate, K_2CO_3 , in 10 parts H_2O . Also made from potassium sulphate, K_2SO_4 , and barium hydroxide, $\text{Ba}(\text{OH})_2$. Purified by dissolving in alcohol. *Liquor potassæ* contains 5 per cent of KOH. Crystals of $\text{KOH} \cdot 2\text{H}_2\text{O}$ are quadratic. Corrodes glass, porcelain and Pt; must be boiled in Ag vessels. Strong base and great absorbent of CO_2 .

Potassium Salts with Oxygen Acids.

Potassium nitrate (saltpetre; KNO_3 ; $\text{KO} - \text{NO}_2$): Occurs in nature, with other nitrates, in India, Egypt, etc. Made from calcium nitrate, $\text{Ca}(\text{NO}_3)_2$, of the saltpetre caves; also, from the animal offal in the nitre beds (Chili saltpetre, NaNO_3); $\text{KCl} + \text{NaNO}_3 = \text{NaCl} + \text{KNO}_3$.

One hundred parts water dissolves

	At 100° C.	At 0° C.
Of NaNO_3	200 parts.	80 parts.
Of KCl	50 parts.	30 parts.
Of KNO_3	250 parts.	13 parts.
Of NaCl	39 parts.	36 parts.

KNO_3 is in anhydrous, rhombic crystals, longitudinal striation; cooling taste; melts at 340°C . *Sal prunelle* is fused saltpeter, molded in balls. Above 340°C , KNO_3 becomes $\text{KNO}_2 + \text{O}$; then $\text{K}_2\text{O} + \text{NO} + \text{O}$. Insoluble in alcohol. Gunpowder is a mechanical mixture of $2\text{KNO}_3 + \text{C}_3 + \text{S}$; when burned $= \text{K}_2\text{S} + 3\text{CO}_2 + \text{N}_2$.

Composition of various gunpowders:

	Parts of KNO_3 .	Parts of C.	Parts of S.
United States and France.....	75	12.5	12.5
Austria.....	76	11.5	12.5
Swiss globular.....	76	14.0	10.0
Sweden.....	76	9.0	15.0
England.....	75	10.0	10.0
Prussia.....	75	13.5	11.5

Potassium nitrite (KNO_2): Made by heating 1 part KNO_3 with 2 parts Pb. Forms small, deliquescent, prismatic crystals. Insoluble in alcohol, but very soluble in water.

Potassium hypochlorite (KClO): Made by passing Cl into dilute solution of KOH, or by decomposing CaO_2Cl_2 . When heated it forms KCl and KClO_3 . Used in bleaching. *Eau de Javelle* contains free HClO.

Potassium chlorate (KClO_3): Made by passing Cl into hot, concentrated solution of KOH or K_2CO_3 ; $6\text{KOH} + 6\text{Cl} = 5\text{KCl} + \text{KClO}_3 + 3\text{H}_2\text{O}$; or Cl passed into solution of KCl and $\text{Ca}(\text{OH})_2$; $6\text{Ca}(\text{OH})_2 + 6\text{Cl}_2 = 5\text{CaCl}_2 + \text{Ca}(\text{ClO}_3)_2 + 6\text{H}_2\text{O}$ and $(\text{CaClO}_3)_2 + 2\text{KCl} = \text{CaCl}_2 + 2\text{KClO}_3$. Is in shining, monoclinic, tabular crystals. Dissolves in 17 parts water at 15°C ., or 2 parts at 100°C . Melts at 331°C . and decomposes at 350°C .; $2\text{KClO}_3 = \text{KCl} + \text{KClO}_4 + \text{O}_2$; at higher temperatures the perchlorate, KClO_4 , yields O and becomes KCl; the addition of manganese dioxide, MnO_2 , facilitates the liberation of O at lower temperatures and more regularly. It is a strong oxidizer; explodes by friction when mixed with S, P, Sb_2S_3 , etc.; when mixed with sugar, H_2SO_4 will explode it. Used in pyrotechnics, matches, *white gunpowder* ($\text{K}_4\text{FeC}_6\text{N}_6$, 28 parts; sugar, 23 parts; KClO_3 , 49 parts); cannon priming ($\text{Sb}_2\text{S}_3 + \text{KClO}_3$), etc. With HCl it yields Cl_2O_4 and Cl.

Potassium bromate (KBrO_3): Made by action of Br on KOH, similar to the formation of KClO_3 . Forms rhombohedral, difficultly soluble crystals. Used for volumetric Br solution.

Potassium perbromate (KBrO_4).

Potassium iodate (KIO_3): Made by action of I on KClO_3 . Forms small, regular crystals, soluble in 13 parts of H_2O at 15°C . Poisonous. Melts at 560°C . Used in organic analysis.

Potassium periodate (KIO_4): When KIO_3 is melted, it forms KIO_4 , O, and KI.

Potassium carbonate: Made from vegetable ashes, cream of tartar, etc.

The *neutral carbonate* $\left\{ \text{K}_2\text{CO}_3; \text{O}=\text{C} \begin{array}{l} \text{OK} \\ \text{OK} \end{array} \right\}$ melts at 834°C .; is volatile in strong white heat. Soluble in 1 part H_2O at 15°C ., or 0.7 part at 100°C .; insoluble in alcohol. Deliquescent, monoclinic crystals ($2\text{K}_2\text{CO}_3 + 3\text{H}_2\text{O}$). Heat expels H_2O , but not CO_2 .

The *acid carbonate* $\left\{ \text{KHCO}_3; \text{O}=\text{C} \begin{array}{l} \text{OH} \\ \text{OK} \end{array} \right\}$, also called *bicarbonate*, is made by passing CO_2 through a solution of K_2CO_3 . Forms monoclinic, anhydrous prisms; soluble in 4 parts of H_2O at 15°C . At 106°C ., and slowly at lower temperatures when in aqueous solution, it loses CO_2 and H_2O .

Potassium sulphates:

K_2SO_4 , neutral (secondary).

KHSO_4 , acid (di—, or primary).

$\text{K}_2\text{S}_2\text{O}_7$, neutral pyrosulphate.

KHS_2O_7 , acid pyrosulphate.

The *neutral sulphate* (secondary, K_2SO_4) is often obtained as a by-product. Forms colorless, hard, anhydrous, rhombic crystals; soluble in 10 parts H_2O at 15°C ., and in 4 parts at 100°C .; insoluble in alcohol.

The *acid sulphate* (di—, bi—, or primary, KHSO_4) crystallizes from excess of H_2SO_4 in hydrated, rhombic tables; after fusion, at 200°C ., it forms anhydrous, monoclinic crystals, easily soluble in H_2O .

The *neutral pyrosulphate* ($\text{K}_2\text{S}_2\text{O}_7$) is formed when KHSO_4 is highly heated; $2\text{KHSO}_4 = \text{K}_2\text{S}_2\text{O}_7 + \text{H}_2\text{O}$. At higher temperatures it splits into SO_3 and K_2SO_4 .

Potassium selenite (K_2SeO_4) and *potassium tellurate* (K_2TeO_4) are similar to the sulphate.

Potassium sulphites: The *neutral sulphite* ($\text{K}_2\text{SO}_3 + 2\text{H}_2\text{O}$) forms deliquescent, monoclinic crystals, which are easily decomposed. The *bisulphite* (acid, KHSO_3) forms large, monoclinic crystals; soluble in H_2O , but insoluble in alcohol.

Potassium phosphates: There are several of them, corresponding to the acids.

The *primary phosphate* (KH_2PO_4) is very soluble.

The *secondary phosphate* (K_2HPO_4) is difficult to crystallize.

The *tertiary phosphate* (neutral, K_3PO_4) forms large quadratic crystals, which are very soluble.

The *neutral pyrophosphate* ($\text{K}_4\text{P}_2\text{O}_7 + 3\text{H}_2\text{O}$) forms a white, deliquescent, crystalline mass, which loses H_2O at 300°C . It is very soluble.

The *acid pyrophosphate* ($\text{K}_2\text{H}_2\text{P}_2\text{O}_7$) is deliquescent.

The *metaphosphate* (KPO_3) is opaque and little soluble.

Potassium arsenates: There are three, corresponding to the acids.

The *primary arsenate* (KH_2AsO_4) forms quadratic crystals.

The *secondary arsenate* (K_2HAsO_4) is not crystalline.

The *tertiary arsenate* (K_3AsO_4) forms deliquescent crystals.

Potassium arsenites: They have been but little studied. The *secondary arsenite* (K_2HASO_3) is very soluble.

Potassium antimonates: The *metantimonate* (KSbO_3) is formed by melting Sb with KNO_3 . From its aqueous solution, the *secondary antimonate* (K_2HSbO_4) results. By melting KSbO_3 with KOH the *pyroantimonate* ($\text{K}_4\text{Sb}_2\text{O}_7$) is formed, which, with boiling H_2O , forms the *acid pyroantimonate* ($\text{K}_2\text{H}_2\text{Sb}_2\text{O}_7$), the reagent for Na.

Potassium borate (KBO_2): Also $\text{K}_2\text{B}_2\text{O}_7 + 5\text{H}_2\text{O}$. Very soluble.

Potassium silicate (water-glass K_4SiO_4): Amorphous. Soluble in H_2O .

Other Potassium Salts.

Potassium sulphides: This metal forms several compounds with S.

The *monosulphide* (K_2S) is made by reducing K_2SO_4 in H, or by C. It is deliquescent, and attracts O from the air, forming $\text{K}_2\text{S}_2\text{O}_3$ and KOH .

The *sulphydrate* (KHS) is formed (also KOH) when K_2S is dissolved in H_2O . Is a colorless, alkaline, deliquescent salt. At 175° to 200°C . loses H_2O , and melts at red-heat to a red mass.

The *polysulphides* (K_2S_2 ; K_2S_4 ; K_2S_5) are formed by melting S with K_2CO_3 . Found in *hepar sulphuris*.

Potassium amide (KH_2N) is formed by passing NH_3 over melted K. Blue while melted, brownish when cold. With H_2O it decomposes into KOH and NH_3 . Burns when heated in O. Heated in an atmosphere free from O it becomes K_3N .

Tests for Potassium.

1. Colors flame violet. Flame appears red when viewed through blue glass or indigo solution.

2. The spectrum has lines in red and violet.

3. Precipitated as acid tartrate ($C_4H_6KO_6$), by tartaric acid ($C_4H_6O_6$), or acid sodium tartrate ($C_4H_5NaO_6$).

4. Forms yellow salt ($K_2Cl_2.PtCl_4$) with platinic chloride ($PtCl_4$).

5. Forms gelatinous silicofluoride (K_2SiF_6), with hydrofluosilicic acid (H_2SiF_6).

6. Precipitated by sodio-cobaltic nitrite, $6NaNO_2.CO_2(NO_2)_6$, forms yellow potassio-cobaltic nitrite, $6KNO_2.CO_2(NO_2)_6$.

7. Most of the salts are soluble in H_2O .

XLIV. RUBIDIUM.

Symbol, Rb. Monad. Atomic weight, 85.5. Specific gravity, 1.52. Melts at $38.5^\circ C$.

Discovered (1860) by Bunsen and Kirchhoff in Dürckheim water, together with Cs, by spectrum analysis.

Occurs in small quantities in mineral springs.

Made by platinic chloride ($PtCl_4$) compound, which is less soluble than the potassium chloride.

Properties are similar to K. Silver-white, slightly yellow metal. Remains soft down to $-10^\circ C$. Distills below red-heat with bluish-green vapor.

Rubidium chloride ($RbCl$): Forms cubes and is more soluble than KCl.

Rubidium bromide ($RbBr$): Forms cubes and octahedra.

Rubidium alum, *platinum double chloride*, and *acid tartrate* are less soluble than corresponding K salts.

Other rubidium salts are similar to corresponding K salts.

Rubidium spectrum: This has two dark-red lines and two less luminous violet ones.

XLV. CÆSIUM.

Symbol, Cs. Monad. Atomic weight, 132.583. Specific gravity, 1.88. Melts at $27^\circ C$.

Discovered with Rb (1860) by Bunsen and Kirchhoff.

Occurs in Dürckheim water (1 ton contains 3 grains $CsCl$); also in Kreuznach mother lye (1:3,000,000 of the dry residue).

Cæsium hydrate ($CsOH$): This salt is strongly alkaline, and corrodes Pt.

Cæsium acid tartrate ($C_4H_5CsO_6$): This is more soluble than the Rb salt.

XLVI. SODIUM.

Symbol, Na. Monad. Atomic weight, 22.998. Specific gravity, 0.972. Melts at $95.6^\circ C$. Boils at 860° to $1,000^\circ C$. *Synonym, natrium.*

Discovered by Sir Humphrey Davy in 1808.

Occurs only in compounds in nature. In sea water as NaCl (2.77 per cent); in New Iberia as NaNO_3 ; in Stassfurth as rock salt; in Chili as saltpetre; in ashes of marine plants, etc.

Made by distillation of the carbonate (Na_2CO_3) with charcoal at white heat.

Na is easily oxidized in air. Withdraws O from water and ice; when water is hot, above 60°C ., it ignites the liberated H, and colors it yellow. Does not unite with Cl or Br, in the cold, but readily on heating.

Sodium hydride (Na_2H): *Formed* when H is absorbed by Na between 300° and 420°C . Is a silver-white metallic mass of 0.959 specific gravity. Soft above and brittle below 0°C .

Halogen Compounds of Sodium.

Sodium chloride (NaCl): *Found* native as rock salt (which is ditherrmanous). *Made* from sea water and saline springs by concentration (to 20 per cent) and crystallization; graduation process or freezing process. Crystals form in cubes, and aggregate. Thirty-six parts dissolve in 100 of H_2O at 0°C . or 39 parts in 100 at 100°C . *Saturated solution* contains 26.5 per cent NaCl, and freezes at -21°C . Is insoluble in alcohol. At -10°C . crystallizes in tabular crystals of $\text{NaCl} + 2\text{H}_2\text{O}$; these lose H_2O at 0°C ., and become anhydrous. Specific gravity is 2.18. Melts at 772°C . Volatile at red heat. Not deliquescent. *Commercial salt* contains impurities (MgCl_2 ; MgSO_4 ; KCl ; Na_2SO_4 ; etc.), some of which are deliquescent.

Sodium bromide (NaBr): When crystallized below 30°C . forms monoclinic prisms containing $2\text{H}_2\text{O}$. Soluble in H_2O . Little soluble in alcohol. Melts at 628°C .

Sodium iodide (NaI): Below 40°C . contains $2\text{H}_2\text{O}$. Easily soluble in H_2O and alcohol. Melts at 628°C .

Sodium silicofluoride (NaSiF_6): Forms difficultly soluble hexagonal crystals. Soluble in 153 parts H_2O at 17°C . and in 40 parts at 100°C .

Sodium fluoride (NaF): Cubes; soluble in 25 parts H_2O .

Sodium cyanide (NaCN): With $2\text{H}_2\text{O}$ or $\frac{1}{2}$ (H_2O) in molecule.

Sodium Oxides and Hydrate.

Sodium monoxide (Na_2O): Is a gray mass.

Sodium dioxide (Na_2O_2): *Made* by passing O over melted Na. White soluble mass. Forms crystals of $\text{Na}_2\text{O}_2 + 8\text{H}_2\text{O}$; also double salt with H_2O_2 ; $\text{Na}_2\text{O}_2 \cdot \text{H}_2\text{O}_2 + 4\text{H}_2\text{O}$.

Sodium hydrate (hydroxide, caustic soda, NaOH): *Made* in various

ways; similar to carbonate; from carbonate and calcium hydrate, $\text{Ca}(\text{OH})_2$. White crystalline, caustic, alkaline. Not deliquescent, but attracts CO_2 from the atmosphere. Dissolves in H_2O with great evolution of heat. Transparent crystals of $2\text{NaOH} + 7\text{H}_2\text{O}$; melts at 6°C .

Sodium Salts with Oxygen Acids.

Sodium nitrate (NaNO_3): Native in Chili. Hexagonal crystals. Melts at 316° to 319°C . Dissolves in 1.25 parts H_2O at 15°C . Deliquesces. Similar to KNO_3 . Used for manufacture of HNO_3 . When heated it forms $\text{NaNO}_2 + \text{O}$.

Sodium nitrite (NaNO_2): In colorless transparent rhombohedra. Less hygroscopic than KNO_2 . Soluble in alcohol.

Sodium chlorate (NaClO_3): Much more soluble than KClO_3 , hence it can not be made by the same process, but by separating chloric acid and saturating with Na_2CO_3 . Forms in large colorless crystals. Soluble in 1 part H_2O at 15°C .

Sodium perchlorate (NaClO_4): Deliquescent and easily soluble.

Sodium bromate (NaBrO_3): Quadratic crystals, soluble in 3 parts H_2O .

Sodium iodate ($\text{NaIO}_3 + \text{H}_2\text{O}$ and $\text{NaIO}_3 + (\text{H}_2\text{O})_5$): Rhombic crystals, soluble in 12 parts H_2O .

Sodium periodate (NaIO_4): Colorless hexagonal crystals, soluble in 12 parts H_2O .

Sodium carbonate: There are two salts formed by carbonic acid.

The *neutral carbonate* (Na_2CO_3) forms crystals of $\text{Na}_2\text{CO}_3 + 10\text{H}_2\text{O}$ from solutions at ordinary temperatures. Effloresces in air. Melts at 50°C . in the water of crystallization. From this is separated $\text{Na}_2\text{CO}_3 + 2\text{H}_2\text{O}$. Rhombic crystals of $\text{Na}_2\text{CO}_3 + 7\text{H}_2\text{O}$ separate from solutions below 30° to 50°C . Made from ashes of marine plants. Native in Nevada, United States; Hungary, Africa, and South America. Manufactured from NaCl by the following processes:

Leblanc's soda ash process:

1. Conversion of NaCl into Na_2SO_4 by H_2SO_4 . The HCl formed is a valuable by-product.

2. Conversion of Na_2SO_4 into Na_2CO_3 by two steps, which are generally united into one; Na_2SO_4 is reduced by C to Na_2S ; Na_2S is reduced to Na_2CO_3 by CaCO_3 .

Over 200,000 tons are made annually in England by this process. Impure *soda ash* contains 48 to 56 per cent of pure Na_2O as carbonate and hydrate; the balance is sulphate, sulphite, chloride, etc. When crystallized with $10\text{H}_2\text{O}$, it forms monoclinic crystals of *sal soda*. When C is used in great excess, no carbonate is formed, but only hydrate.

Kryolite process:

Kryolite ($6\text{NaF} \cdot \text{Al}_2\text{F}_6$) is heated with CaO , lixiviated, and CO_2 passed through the solution.

Ammonium bicarbonate process:

A concentrated solution of NaCl in H_2O is saturated—first, with ammonia gas, then with CO_2 . Sodium acid carbonate (NaHCO_3), soluble with difficulty, separates and is converted by heat into Na_2CO_3 and CO_2 and H_2O : $2\text{NaHCO}_3 = \text{Na}_2\text{CO}_3 + \text{CO}_2 + \text{H}_2\text{O}$. The NH_3 is converted into NH_4Cl , from which ammonia gas is regained for the next operation by treating with calcium hydrate $\text{Ca}(\text{OH})_2$.

Sodium bicarbonate (or *dicarbonate*, NaHCO_3): Anhydrous monoclinic crystals. Soluble in 12 parts cold H_2O . Loses CO_2 by heating.

Sodium sesquicarbonate ($\text{Na}_2\text{CO}_3 + 2\text{NHCO}_3 + 2\text{H}_2\text{O}$); found native.

Water, at 0°C . dissolves 18 per cent of Na_2CO_3 ; at 38°C ., 138 per cent; at 104°C ., 121 per cent.

Sodium sulphate: The neutral ($\text{Na}_2\text{SO}_4 + 10\text{H}_2\text{O}$) is in monoclinic crystals, which melt in the water of crystallization at 33°C . The anhydrous salt melts at 861°C . Another hydration is $\text{Na}_2\text{SO}_4 + 7\text{H}_2\text{O}$; it forms rhombic crystals.

Water at 0°C . dissolves 12 per cent of Na_2SO_4 ; at 18° , 48 per cent; at 33° , 322 per cent; at 50° , 263 per cent; at 100° , 238 per cent.

Sodium acid (di—or bi—) *sulphate* ($\text{NaHSO}_4 + \text{H}_2\text{O}$) forms in triclinic crystals. Anhydrous at 50°C .

Sodium and potassium sulphate (NaKSO_4): Forms in hexagonal crystals.

Sodium sulphite: Monoclinic $\text{Na}_2\text{SO}_3 + 7\text{H}_2\text{O}$, or prismatic NaHSO_3 , both soluble, and used as antizymotics.

Sodium hyposulphite (thiosulphate, $\text{Na}_2\text{S}_2\text{O}_3 + 5\text{H}_2\text{O}$): Large, hygroscopic, readily soluble, monoclinic crystals. Made from residue of soda manufacture. Great reductent. Converts I into HI. Dissolves halogen salts of Ag. Used in photography and volumetric analysis.

Sodium phosphates: There are several of them.

The *primary* ($\text{NaH}_2\text{PO}_4 + \text{H}_2\text{O}$) forms rhombic crystals.

The *secondary* ($\text{Na}_2\text{HPO}_4 + 12\text{H}_2\text{O}$) forms monoclinic crystals.

The *tertiary* ($\text{Na}_3\text{PO}_4 + 12\text{H}_2\text{O}$) forms hexagonal crystals.

The most important is $\text{Na}_2\text{HPO}_4 + 12\text{H}_2\text{O}$. It is formed from the tertiary by CO_2 of the air. Made from bone ashes, which are treated with H_2SO_4 , and then saturated with Na_2CO_3 . Has feeble alkaline reaction. Soluble in 4 parts cold or 2 parts hot water. When crystallized at 80°C ., $\text{Na}_2\text{HPO}_4 + 7\text{H}_2\text{O}$ is formed. *Microcosmic salt* is $\text{NaNH}_4\text{HPO}_4 + 4\text{H}_2\text{O}$. A double salt with K is also formed; NaKHPO_4 .

The *pyrophosphate* is either $\text{Na}_4\text{P}_2\text{O}_7 + 10\text{H}_2\text{O}$ or $\text{Na}_2\text{H}_2\text{P}_2\text{O}_7$. They form a yellow precipitate with AgNO_3 .

The *metaphosphate* (NaPO_3) is in several polymeric modifications. Made by heating microcosmic salt.

Sodium arsenate ($\text{Na}_3\text{AsO}_4 + 12\text{H}_2\text{O}$, or with $7\text{H}_2\text{O}$): The official *liquor sodii arseniatis* contains 1 per cent of $\text{Na}_2\text{HAsO}_4 + 7\text{H}_2\text{O}$. This, when heated, furnishes the pyroarsenate ($\text{Na}_4\text{As}_2\text{O}_7$). There is also $\text{NaH}_2\text{AsO}_4 + \text{H}_2\text{O}$.

Sodium antimonate (NaH_2SbO_4): When first precipitated by KH_2SbO_4 it is not entirely insoluble, but soon becomes granular and totally insoluble. The *meta*— and *pyro*— salts are also known.

Sodium borate: The *neutral* ($\text{NaBO}_2 + 4\text{H}_2\text{O}$) is triclinic and the *diborate* ($\text{Na}_4\text{B}_4\text{O}_7 + 10\text{H}_2\text{O}$) is monoclinic. Native in East India and California. Soluble in 14 parts cold, or with $\frac{1}{2}$ part hot water. Melts at red heat. Dissolves metallic oxides. Has alkaline reaction.

Sodium silicate (water glass, $\text{Na}_2\text{SiO}_3 + 8\text{H}_2\text{O}$) is monoclinic. Used as a fire-proof varnish, in surgical dressing, paste, in soaps, etc.

Sodium sulphides: The *monosulphide* (Na_2S) and *sulphydrate* (NaHS), also *polysulphides*.

Sodium amide (NaNH_2): In greenish-blue crystals.

Tests for Sodium.

1. Yellow color of flame.
2. Two yellow lines in spectrum.
3. Precipitated from alkaline solution by potassium acid pyroantimonate as $\text{Na}_2\text{H}_2\text{Sb}_2\text{O}_7$.
4. Precipitated by potassium stannoso-chloride.

XLVII. LITHIUM.

Symbol, Li. Monad. Atomic weight, 7.007. Specific gravity, 0.5936. Melts at 180°C .

Discovered in combination by Arfoedson in 1817; the metal by Brande in 1822.

Occurs in lepidolite, triphyllin, spodumen, in soils, tobacco ashes, etc.

Silver white, soft metal. Decomposes H_2O without ignition of H. Burns when heated in air. Colors flame crimson.

Lithium chloride (LiCl): Octohedral crystals of $\text{LiCl} + 2\text{H}_2\text{O}$ below 10°C . Melts at 598°C . Deliquescent. Soluble in alcohol, also in mixture of alcohol and ether.

Lithium bromide (LiBr): Deliquescent.

Lithium iodide ($\text{LiI} + 3\text{H}_2\text{O}$): Deliquescent. Melts at 446°C .

Lithium hydrate (LiOH): Similar to other alkaline hydrates.

Lithium nitrate (LiNO_3): Anhydrous rhombohedral crystals at 15°C . Below 10°C . forms thin deliquescent prisms of $2\text{LiNO}_3 + 5\text{H}_2\text{O}$.

Lithium carbonate (Li_2CO_3): Precipitated by K_2CO_3 , etc., from other

Li salts. More soluble in cold than in hot H_2O . The *bicarbonate* (LiHCO_3) occurs in mineral springs.

Lithium sulphate ($\text{Li}_2\text{SO}_4 + \text{H}_2\text{O}$): Easily soluble, monoclinic crystals.

Lithium phosphate ($2\text{Li}_3\text{PO}_4 + \text{H}_2\text{O}$): White crystalline powder. Soluble in 2,500 parts cold H_2O ; still less in ammonia water. Made by mixing LiCl with Na_2HPO_4 .

The *acid phosphate* (LiH_2PO_4) is made by saturating the neutral phosphate with H_3PO_4 . Forms large deliquescent crystals.

Tests for lithium: Gives carmine color to flame. Spectrum has red, orange, and blue lines. Carbonate and neutral phosphate are difficultly soluble.

XLVIII. AMMONIUM.

Formula, NH_4 . Monatomic group. Molecular weight, 18.

The group NH_4 has not been isolated. Is isomorphous and properties are similar to the alkali metals. All salts yield NH_3 when treated with KOH or NaOH .

Ammonium amalgam: Made from Na amalgam and NH_4Cl , or by electrolysis of NH_4Cl in presence of Hg . Forms a voluminous, butyraceous, unstable mass.

Ammonium chloride (sal ammoniac, NH_4Cl): Formed by direct union of NH_3 and HCl . Made from impure carbonate by saturating with HCl ; also by subliming from gas water, animal offal, etc. Forms regular octohedral crystals, often in peculiar feathery aggregation. Soluble in $2\frac{1}{2}$ parts H_2O at 15°C ., or 1 part at 100°C .; insoluble in absolute alcohol. Is volatile, and dissociates on heating crystals or solution; reunites when cooled.

Ammonium bromide (NH_4Br): Forms in colorless prisms. Used in photography.

Ammonium iodide (NH_4I): Made by mixing $(\text{NH}_4)_2\text{SO}_4$ and KI . Forms hygroscopic cubes. Very soluble in H_2O and alcohol. Forms double compounds with metallic iodides. Dissolves I .

Ammonium fluoride (NH_4F): Forms in small deliquescent prisms.

Ammonium cyanide (NH_4CN): Made by passing NH_3 over red hot charcoal, or by subliming KCN with NH_4Cl . Forms colorless cubes.

Ammonium hydroxide (NH_4OH): Unknown in free state, but is supposed to exist in aqueous solutions.

Ammonium nitrate (NH_4NO_3): Made by neutralizing HNO_3 with NH_3 or $(\text{NH}_4)_2\text{CO}_3$; on large scale, by decomposition of $(\text{NH}_4)_2\text{SO}_4$ and KNO_3 . Forms six-sided rhombic, deliquescent crystals. Dissolves in H_2O with great refrigeration. Melts at 168°C . Sublimes at 200°C . Above 210°C . decomposes into N_2O and H_2O . Absorbs dry NH_3 , forming compounds.

Ammonium nitrite (NH_4NO_2): Formed by electric discharges in moist air, also by ozone and dilute solution of NH_3 . Forms soluble crystals. Easily decomposed into N_2 and $2\text{H}_2\text{O}$. Detonates at 60° to 70° C. and by percussion.

Ammonium chlorate (NH_4ClO_3): Very explosive, spontaneously, and at 100° C.

Ammonium carbonate: There are three known.

The commercial sesquicarbonate $[\text{NH}_4\text{CO}_2(\text{NH}_2) + 2(\text{NH}_4\text{HCO}_3)]$ is a mixture of carbamate and bicarbonate. Made by subliming NH_4Cl with CaCO_3 . Converted by H_2O into carbonate and bicarbonate; $\text{NH}_4\text{CO}_2(\text{NH}_2) + 2(\text{NH}_4\text{HCO}_3) + \text{H}_2\text{O} = (\text{NH}_4)_2\text{CO}_3 + 2(\text{NH}_4\text{HCO}_3)$.

The neutral carbonate $[(\text{NH}_4)_2\text{CO}_3]$ is made by adding NH_3 to a solution of the commercial carbonate. Is unstable and loses NH_3 .

The bicarbonate (NH_4HCO_3) forms hard rhombic prisms, soluble in 8 parts H_2O . Loses CO_2 on standing.

Ammonium (neutral) sulphate $[(\text{NH}_4)_2\text{SO}_4]$: Forms colorless rhombic crystals. Soluble in 2 parts of cold and 1 part of boiling H_2O . Melts at 140° C.; at higher temperatures forms $\text{NH}_3 \cdot \text{N} \cdot \text{H}_2\text{O}$ and $(\text{NH}_4)_2\text{SO}_3$.

The acid sulphate $[(\text{NH}_4)\text{HSO}_4]$ forms rhombic crystals.

The ammonium potassium sulphate $[(\text{NH}_4)\text{KSO}_4]$ forms colorless crystals.

Ammonium sulphite $[(\text{NH}_4)_2\text{SO}_3]$: Made by saturating an aqueous solution of SO_2 with NH_3 and precipitating with alcohol. Forms monoclinic crystals.

Ammonium phosphates: There are several of them.

The primary $[(\text{NH}_4)\text{H}_2\text{PO}_4]$ forms quadratic crystals.

The secondary $[(\text{NH}_4)_2\text{HPO}_4]$ formed when the tertiary loses NH_3 . Monoclinic crystals.

The tertiary $[(\text{NH}_4)_3\text{PO}_4]$ forms a crystalline mass.

The ammonium, sodium, hydrogen phosphate $[(\text{NH}_4)\text{NaHPO}_4]$ is the microcosmic salt.

Ammonium arsenates: Three salts are known similar to the phosphates.

Ammonium (pyro) borate $[(\text{NH}_4)_2\text{B}_4\text{O}_7 + 4\text{H}_2\text{O}]$ forms quadratic crystals. Another borate $[(\text{NH}_4)\text{B}_5\text{O}_8]$ forms rhombic crystals.

Ammonium sulphide $[(\text{NH}_4)_2\text{S}]$: Not known in isolation, but in solution (NH_4HS exists free). The aqueous solution is used in analysis for precipitating the Fe group.

Ammonium sulphhydrate (NH_4HS): Forms large colorless tables. Made by saturating an alcoholic solution of NH_3 with H_2S . The aqueous solution is used as a reagent. It soon becomes yellow by the formation of polysulphides and polythionates.

Tests for ammonium salts: 1. Add caustic alkalis to liberate NH_3

and bring HCl in contact with the gas. It forms white fumes of NH_4Cl .

2. *Nessler's reagent* is an alkaline solution of $\text{HgI}_2 \cdot 2\text{KI}$, and forms brown precipitate with NH_3 salts.

3. *Tartaric acid* forms a precipitate.

4. *Platinic chloride* forms a precipitate.

5. *Sodiocobaltic nitrite* forms a precipitate.

6. *Characteristic odor* when NH_3 is liberated by *caustic alkalies*.

XLIX.—HYDROXYLAMINE.

Formula, NH_2O ; NH_2OH .

This compound is intermediate in composition between NH_3 and NH_4OH . Formed by the reduction of HNO_3 , or HNO_2 , through nascent hydrogen; $\text{HNO}_3 + 3\text{H}_2 = \text{NH}_2\text{OH} + 2\text{H}_2\text{O}$.

Hydroxylamine salts are similar to NH_3 salts. Very unstable. Used as reducents in photography.

L.—METALS OF THE ALKALINE EARTHS.

	Ca.	Sr.	Ba.
Atomic weight.....	39.99	87.374	136.763
Specific gravity.....	1.58	2.54	4.3
Melting point.....	500° C.	490° C.	500° C. ?
Discovered.....	1807	1807	1808
By	Davy.	Davy.	Davy.

Characteristics.

All are diads.

All decompose cold water.

With halogens form compounds; $M^{\text{II}}\text{Cl}_2$, etc.

With oxygen form compounds; $M^{\text{II}}\text{O}$ and $M^{\text{II}}\text{O}_2$.

Form hydroxides; $M^{\text{II}}(\text{OH})_2$.

Strong bases (Ba, strongest; Ca, weakest).

Sulphates are insoluble or little soluble.

Carbonates are insoluble but the bicarbonates are soluble.

Hydrates are soluble (alkaline reaction).

LI.—BARIUM.

Symbol, Ba. Diad. Atomic weight, 136.763. Specific gravity, 4.3. Melts, 475° C.

Occurs only in salts, especially as sulphate (*heavy spar*) and carbonate (*witherite*).

Prepared by electrolysis of the fused BaCl_2 , either as amalgam or direct.

Is a bright yellow metal. Can not be distilled. Rapidly oxidizes in the air.

Barium chloride ($\text{BaCl}_2 + 2\text{H}_2\text{O}$): Forms colorless, rhombic, poisonous plates. Loses water of crystallization at 100°C . Melts at red heat. Soluble in 2 parts cold or $1\frac{1}{2}$ parts hot water. Insoluble in absolute alcohol and in HCl . Continued heat in air forms a small percentage of BaO .

Barium bromide ($\text{BaBr}_2 + 2\text{H}_2\text{O}$): Easily soluble in alcohol.

Barium iodide ($\text{BaI}_2 + 7\text{H}_2\text{O}$): Similar to BaBr_2 . When heated in air forms BaO and I_2 .

Barium fluoride (BaF_2): Difficultly soluble.

Barium silicofluoride (BaSiF_6): Requires 3,500 parts of H_2O for solution at 15°C ., or 1,200 at 100°C . Employed as a reagent to separate Ba from Sr .

Barium cyanide [$\text{Ba}(\text{CN})_2$]: Formed when BaO and C are heated in current of air. White, deliquescent crystals, easily decomposed by CO_2 .

Barium monoxide (BaO): Made by heating $\text{Ba}(\text{NO}_3)_2$ at high heat. BaCO_3 is not easily converted in BaO and CO_2 . Gray amorphous mass. Specific gravity is 5.3 to 5.7.

Barium dioxide (BaO_2 ; also $\text{BaO}_2 + 8\text{H}_2\text{O}$): Made by passing O over BaO at low red heat. White, insoluble powder. Used for making H_2O_2 .

Barium hydrate [$\text{Ba}(\text{OH})_2$; also $\text{Ba}(\text{OH})_2 + 8\text{H}_2\text{O}$]: Made by addition of BaO to H_2O , the reaction causing heat. Precipitated from Ba salts by NH_3 . Soluble in 20 parts H_2O at 15°C ., or 3 at 100°C . Strongly alkaline. Absorbs CO_2 .

Barium nitrate [$\text{Ba}(\text{NO}_3)_2$]: Forms colorless octohedra. Soluble in 12 parts H_2O at 15°C ., or 3.5 at 100°C . Used in pyrotechnics for green fire.

Barium chlorate [$\text{Ba}(\text{ClO}_3)_2 + \text{H}_2\text{O}$]: Soluble in 4 parts of H_2O at 15°C .

Barium carbonate (BaCO_3): Native as *witherrite*, rhombic crystals; artificial as a white powder. Insoluble in H_2O . Soluble in H_2O containing CO_2 , [$\text{BaH}_2(\text{CO}_3)_2$]. Melts at white heat, losing some CO_2 .

Barium sulphate (BaSO_4): Native as *heavy spar*, rhombic crystals. Precipitated from Ba salts by H_2SO_4 , or soluble sulphates as a heavy white powder. Totally insoluble in H_2O and all acids, except concentrated H_2SO_4 at 100°C . The *disulphate* [$\text{BaH}_2(\text{SO}_4)_2$] is formed when freshly precipitated; it is also known with $2\text{H}_2\text{O}$. Water at once decomposes this salt. The artificial salt is used as *blanc fix* for paint; the finely ground native salt is used to adulterate white lead.

Barium sulphite (BaSO_3): A white crystalline powder, almost insoluble in H_2O ; soluble in dilute H_2SO_4 .

Barium hyposulphate (BaS_2O_6): Large rhombic (with H_2O) or monoclinic (with $4\text{H}_2\text{O}$) crystals.

Barium hyposulphite ($\text{BaS}_2\text{O}_3 + \text{H}_2\text{O}$): Difficultly soluble crystals.

Barium phosphates:

The *primary* [$\text{BaH}_4(\text{PO}_4)_2$] is soluble in little but decomposed by much H_2O . Forms colorless crystals with acid reaction.

The *secondary* (BaHPO_4) is insoluble in H_2O , but soluble in dilute HNO_3 or HCl . Forms a white crystalline powder.

The *tertiary* [$\text{Ba}_3(\text{PO}_4)_2$] is insoluble in H_2O , but soluble in dilute HNO_3 or HCl .

Barium silicate ($\text{BaSiO}_3 + 6\text{H}_2\text{O}$): Forms as insoluble rhombic crystals in bottles containing baryta water.

Barium sulphide (BaS): A gray mass when pure. Decomposes in air. With H_2O it splits into $\text{Ba}(\text{OH})_2$ and $\text{Ba}(\text{SH})_2$.

Barium hydrosulphide [$\text{Ba}(\text{SH})_2$]: Very soluble colorless crystals.

Tests for barium salts:

1. Green color of flame.
2. Many green lines in spectrum.
3. Precipitated as insoluble BaSO_4 by H_2SO_4 or soluble sulphates.
4. Precipitated as BaCrO_4 (insoluble in acetic acid) by chromates. Precipitated by bichromates.
5. Precipitated as BaCO_3 by soluble carbonates.
6. The silicofluoride is little soluble.

LII. STRONTIUM.

Symbol, Sr. Diad. Atomic weight, 86.37; specific gravity, 2.542. Occurs native as *celestine* (SrSO_4) and *strontianite* (SrCO_3), and with lime stones. Is a bright yellow ductile metal. Decomposes H_2O . Oxidizes in air.

Strontium chloride ($\text{SrCl}_2 + 6\text{H}_2\text{O}$): Forms hexagonal deliquescent soluble prisms. The anhydrous melts at 825°C . Burns with red flame. Soluble in alcohol.

Strontium bromide ($\text{SrBr}_2 + 6\text{H}_2\text{O}$): Very soluble hexagonal crystals.

Strontium iodide ($\text{SrI}_2 + 7\text{H}_2\text{O}$): Very soluble hexagonal plates.

Strontium fluoride (SrF_2): Little soluble.

Strontium silicofluoride ($\text{SrSiF}_6 + 2\text{H}_2\text{O}$): Very soluble monoclinic crystals.

Strontium oxide (SrO): A white mass similar to BaO . Made by igniting $\text{Sr}(\text{NO}_3)_2$.

Strontium dioxide ($\text{SrO}_2 + 8\text{H}_2\text{O}$): White powder. At 130°C . it only loses H_2O , but is decomposed by high heat.

Strontium hydrate [$\text{Sr}(\text{OH})_2$]: Similar to $\text{Ba}(\text{OH})_2$ but less soluble.

Strontium nitrate $[\text{Sr}(\text{NO}_3)_2]$: Anhydrous octohedra; crystallize from hot H_2O , and monoclinic $\text{Sr}(\text{NO}_3)_2 + 4\text{H}_2\text{O}$ from cold H_2O . Soluble in 5 parts H_2O at 15°C ., or $\frac{1}{2}$ at 100°C . Insoluble in alcohol. Used for red fire in pyrotechnics.

Strontium chlorate $[\text{Sr}(\text{ClO}_3)_2 + 5\text{H}_2\text{O}]$: Very soluble and deliquescent.

Strontium carbonate (SrCO_3): Native as rhombic *strontianite*. Artificial as insoluble white powder.

Strontium sulphate (SrSO_4): Native as rhombic *celestine*. Artificial as white crystalline powder. Slightly soluble in H_2O , or concentrated H_2SO_4 . Aqueous solution is precipitated by Ba salts.

The other Sr salts are similar to the corresponding Ba salts.

Tests for strontium:

1. Colors flame red.
2. Red, orange, and blue lines in spectrum.
3. Precipitated by H_2SO_4 ; more soluble than BaSO_4 .
4. Not precipitated by bichromates but by chromates, as SrCrO_4 ; soluble in acetic acid.
5. Not precipitated by H_2SiF_6 .

LIII. CALCIUM.

Symbol, Ca. Diad. Atomic weight, 39.99. Specific gravity, 1.577. Melts, red heat.

Occurs native as: carbonate (limestone, marble, chalk); sulphate (gypsum, alabaster); phosphate (phosphorite, apatite); silicate; fluoride (fluorspar). These compounds form from 6 to 7 per cent of the earth's crust.

Made by electrolysis, etc., like Ba. Brittle white metal. Stable in dry, but oxidizes in moist air. Burns with yellow light. Decomposes H_2O , also diluted HNO_3 , but not concentrated HNO_3 .

Calcium chloride ($\text{CaCl}_2 + 6\text{H}_2\text{O}$): Obtained as a by-product of many manufactures, also by saturating HCl with CaCO_3 . Large transparent hexagonal crystals, melting in their crystal H_2O at 29°C . Soluble in alcohol. Absorbs NH_3 forming $\text{CaCl}_2 + 8\text{NH}_3$. The *anhydrous* melts at 719°C ., losing some Cl as HCl and forming CaO. This melted salt is used as a dehydrator.

Calcium bromide (CaBr_2): White deliquescent salt.

Calcium iodide (CaI_2): White deliquescent salt. Loses all its I by heating.

Calcium fluoride (CaF_2 , fluorspar): Cubes with octohedral cleavage. *Occurs native.* Found in enamel of teeth, in bone, ashes, etc. Almost insoluble in H_2O but somewhat soluble in diluted acids. Used as flux. For etching, and making HF.

Calcium silicofluoride (CaSiF_6): Easily soluble in H_2O .

Calcium oxide (CaO): Made on large scale by burning limestone. White, infusible, giving high light when heated. Used in Drummond's calcium light.

Calcium dioxide ($\text{CaO}_2 + 8\text{H}_2\text{O}$): Similar to BaO_2 and SrO_2 .

Calcium hydroxide [$\text{Ca}(\text{OH})_2$]: Formed when CaO takes up H_2O (evolving great heat). White amorphous mass, soluble in 700 to 800 parts cold H_2O , or 1,700 of boiling H_2O . Used for mortar and cement. Hydraulic cement contains silicate. Lime water attracts CO_2 from the air.

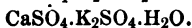
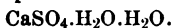
Calcium nitrate [$\text{Ca}(\text{NO}_3)_2$]: Forms in caves, stables, etc. Soluble in H_2O or alcohol. $\text{Ca}(\text{NO}_3)_2 + 4\text{H}_2\text{O}$ is monoclinic.

Calcium hypochlorite [$\text{Ca}(\text{ClO})_2$]: In admixture (or combination) with CaCl_2 it forms *bleaching powder*, from which HCl or H_2SO_4 evolves twice as much Cl as exists in ClO , so that the CaCl_2 must co-operate

in the evolution, hence the formula: $\text{Ca} \begin{array}{c} \text{OCl} \\ \text{Cl} \end{array}$. Used as a disinfectant; for bleaching; making chloroform; etc. Seldom contains more than 38 per cent of active Cl . This is determined by sodium arsenite solution and I. Pure $\text{Ca}(\text{ClO})_2 + 4\text{H}_2\text{O}$ is formed by evaporating solution in vacuo.

Calcium carbonate (CaCO_3): Native or *calcspar*, *arroganite*, *limestone*, *marble*, *chalk*, etc. Insoluble in H_2O unless CO_2 is present to form the *bicarbonate* [$\text{CaH}_2(\text{CO}_3)_2$]. Hardness of H_2O caused by the latter salt. Soap used as a test. Clark's method of softening H_2O . Formation of *stalactites* and *stalagmites* in caves. Kettle stone caused by hard H_2O . Heat expels CO_2 .

Calcium sulphate (CaSO_4): Found with $2\text{H}_2\text{O}$ as *gypsum* and *selenite* in monoclinic crystals, or granular *asabaster*, etc. Little soluble in H_2O ; only one part in 415 at 0°C ., in 398 at 15° , in 359 at 36° , and in 455 at 100° . The anhydrous is known as *calcium anhydride*. The permanent hardness of H_2O may be caused by CaSO_4 . At 200°C . gypsum loses its water of crystallization. When this is only partially expelled, *plaster of paris* is formed which, on mixing with H_2O , again combines and hardens; but when once anhydrous the H_2O is not resumed. Same action in *stucco*, *Keene's cement*, *bandages*, etc. The best result is obtained by heating to 133°C ., when $\frac{3}{4}$ of the $2\text{H}_2\text{O}$ is expelled. The powder when heated "boils" like a liquid until this point is reached. Water containing CaSO_4 and organic matter frequently smells of H_2S from reduction of the sulphate. Salts are formed in which the one H_2O is replaced by some alkaline sulphate:



$\text{CaSO}_4 \cdot \text{Na}_2\text{SO}_4 \cdot \text{H}_2\text{O}$.

Calcium sulphite ($\text{CaSO}_3 + \text{H}_2\text{O}$): Difficult to dissolve in H_2O , but easily dissolved in diluted H_2SO_3 .

Calcium phosphates:

The *primary* or *acid* [$\text{CaH}_4(\text{PO}_4)_2 + \text{H}_2\text{O}$] is made by evaporating the *neutral* with acid. Colorless tables. Decomposes by boiling *neutral* salt and free acid.

The *secondary* ($\text{CaHPO}_4 + 2\text{H}_2\text{O}$) is found native as *monoclinic brushite* in guano.

The *tertiary* or *neutral* [$\text{Ca}_3(\text{PO}_4)_2$] is found in *apatite* (with CaF_2 and CaCl_2), in *phosphorite*, in soils, in bones (85 per cent in bone ashes), in guano. Almost insoluble in H_2O unless Cl_2 or salts are present. Easily soluble in HCl ; HNO_3 or $\text{HC}_2\text{H}_3\text{O}_2$.

Calcium arsenate [$\text{Ca}_3(\text{AsO}_4)_2$]: It is found native as *hardingerite* ($\text{CaHAsO}_4 + \text{H}_2\text{O}$) and in *pharmakolite* ($2\text{CaHAsO}_4 + 5\text{H}_2\text{O}$).

Calcium borate [$\text{Ca}(\text{BO}_2)_2$]: Native as *rhodizite* in regular crystals. *Boro-natro calcite* [$2(\text{CaB}_4\text{O}_7 + \text{NaBO}_2) + 15 \text{H}_2\text{O}$] in Chili.

Calcium silicate (CaSiO_3): Occurs as *Wollastonite* and in compound minerals (native with borate as *datolite*). *Glass*: Mostly silicate of Ca and an alkali. Common glass is made with Na; Bohemian glass is less fusible and made mostly with K; there are also often Al, Fe, and other metals present. Flint glass contains silicate of lead; for dense optical glass lead borate is also used. Colors are given by metallic oxides; red by Au and by cuprous oxide; green by Cr, Fe, Cu; blue by Co; yellow by Ur (fluoresces green); amethyst, by Mn.

Calcium sulphide (CaS) and *calcium hydrosulphide* [$\text{Ca}(\text{SH})_2$]: By boiling $\text{Ca}(\text{OH})_2$ with S a very complicated solution results of sulphides, polysulphides, and hyposulphides, from which S is precipitated by HCl . The *oxysulphides* form orange and yellow crystals.

Phosphorcalcium (Ca_2P_2): Made by heating CaO in vapor of P. A red-brown mass, evolving PH_3 with H_2O .

Calciumsilicum: A gray mass with metallic luster. Decomposed by air or H_2O .

Calciumcarbon: Forms with H_2O acetylene and $\text{Ca}(\text{OH})_2$.

Tests for calcium salts:

1. Colors flame yellow.
2. Yellow, orange, and green lines in spectrum.
3. Precipitated as $\text{Ca}(\text{OH})_2$ by KOH .
4. Oxalic acid precipitates oxalate of calcium, which is insoluble in acetic acid.

LIV.—METALS OF THE ZINC GROUP.

	Mg.	Zn.	Cd.
Atomic weight.....	23.959	64.905	111.77
Specific gravity.....	1.750	7.200	8.60
Melting point.....	Red heat.	412° C.	315-5° C.
Boiling point.....		940° C.	763° C.

They all are diads (MgCl_2 , MgO , ZnCl_2 , etc.). All form insoluble carbonates. The hydrates are precipitated by KOH , NaOH , or NH_4OH .

LV.—MAGNESIUM.

Symbol, Mg . Diad. Atomic weight, 23.959. Specific gravity, 1.75.

Occurs in nature only in compounds. Found as carbonate and silicate in steatite, meerschaum, etc. First isolated (impure) by Davy, next (pure) by Liebig & Bussy. Made by fusing anhydrous MgCl_2 with Na . A white metal. Ductile when hot. Melts at red heat. Sublimes in hexagonal prisms. Burns with great brilliancy when heated in air: The light is rich in ultraviolet rays, hence used in photography. Does not decompose cold, but slowly, boiling H_2O . Dissolves in acids or ammonia salts, liberating H . Not soluble in KOH .

Magnesium chloride (MgCl_2): Occurs in sea water and mineral springs. Made from solution it forms $\text{MgCl}_2 + 6\text{H}_2\text{O}$, which can not be made anhydrous, but forms HCl and MgO and oxychlorides. The anhydrous is made from double ammonia salts; colorless, deliquescent, soluble in alcohol, melts at 708°C ., volatile. $\text{MgCl}_2 + \text{KCl} + 6\text{H}_2\text{O}$ as rhombic crystals. $\text{MgCl}_2 + \text{NH}_4\text{Cl} + 6\text{H}_2\text{O}$ as rhombic crystals. $2\text{MgCl}_2 + \text{CaCl}_2 + 12\text{H}_2\text{O}$ is deliquescent.

Magnesium bromide ($\text{MgBr}_2 + 6\text{H}_2\text{O}$): Properties similar to the chloride.

Magnesium iodide (MgI_2): Deliquescent.

Magnesium fluoride (MgF_2): Native as quadratic *sellaite*. Forms as $\text{MgF}_2 + 6\text{H}_2\text{O}$.

Magnesium silicofluoride (MgSiF_6): Hexagonal.

Magnesium oxide (*Magnesia usta*, MgO): Native as rare *periclase*. Made by heating MgCO_3 or Mg(OH)_2 . Heavy and light *magnesia usta* have the same chemical composition, but the former is made from the heavy carbonate and the latter from the light salt. Soluble in 514 parts H_2O at 15°C ., or 36,000 at 100°C .

Magnesium hydrate [Mg(OH)_2]: Forms as a gelatinous mass when MgO is moistened with 15 parts H_2O . White powder. Insoluble in H_2O , KOH , NaOH , or NH_4OH ; soluble in NH_4Cl and other ammonia salts. Attracts CO_2 from the atmosphere. Native as *brucite*.

Magnesium nitrate [$\text{Mg(NO}_3)_2 + 6\text{H}_2\text{O}$]: Monoclinic, deliquescent, easily soluble in H_2O or alcohol. Melts at 90°C . in water of crystallization.

Magnesium carbonate (MgCO_3): Native as *magnesite* and *hydromagnesite*. Artificial as a basic salt (*magnesia alba*) $3\text{MgCO}_3 + \text{Mg(OH)}_2 + 4\text{H}_2\text{O}$. From the bicarbonate [$\text{Mg}_2\text{H}_2(\text{CO}_3)_2$] the carbonate deposits. Heating expels CO_2 . Forms double salts; $\text{MgK}_2(\text{CO}_3)_2 + 4\text{H}_2\text{O}$, also $\text{MgKH}(\text{CO}_3)_2 + 4\text{H}_2\text{O}$, also with NH_4 and Ca .

Magnesium sulphate (epsom salt, $\text{MgSO}_4 + 7\text{H}_2\text{O}$): Native *epsomite* with $7\text{H}_2\text{O}$ and *krieserite* with $1\text{H}_2\text{O}$. Made from *dolomite* with $7\text{H}_2\text{O}$ of which one differs (lost at 200°C.) from the other six (lost at 150°C.) by being substituted by alkaline sulphates. Soluble in $\frac{4}{5}$ part H_2O at 15°C. ; insoluble in alcohol. Monoclinic.

Magnesium sulphite ($\text{MgSO}_3 + 6\text{H}_2\text{O}$): Hexagonal, slightly soluble crystals.

Magnesium phosphate [neutral, $\text{Mg}_3(\text{PO}_4)_2$]: Made by precipitation. Native as *wagnerite* (pleuroklas) with MgF_2 . The secondary salt (MgHPO_4) is made by precipitation with Na_2HPO_4 . No *acid* salt forms. Double salts form with Na, etc.

Magnesium ammonium phosphate (triple phosphate, $\text{MgNH}_4\text{PO}_4 + 6\text{H}_2\text{O}$): Occurs in urinary deposits, calculi, guano, etc. Insoluble in H_2O or NH_4OH .

Magnesium pyrophosphate ($\text{Mg}_2\text{P}_2\text{O}_7$): Made by heating triple phosphate. Used for quantitative determinations.

Magnesium arsenate [$\text{Mg}_3(\text{AsO}_4)_2$ and $\text{MgHAsO}_4 + 7\text{H}_2\text{O}$.] White, little soluble in H_2O but soluble in acids. $\text{MgH}_4(\text{AsO}_4)_2$ is soluble in H_2O but does not crystallize.

Magnesium ammonium arsenate [$\text{Mg}(\text{NH}_4)\text{AsO}_4 + 6\text{H}_2\text{O}$]: Similar to the phosphate.

Magnesium borate [$\text{Mg}_3(\text{BO}_3)_2 + 9\text{H}_2\text{O}$]: Formed by boiling a mixture of solutions of borax and Mg salt. It redissolves on cooling. Also $\text{Mg}_3\text{B}_3\text{O}_{15}$ as fine needles. This combined with MgCl_2 forms *boricite* and *stassfurtite*.

Magnesium silicate (Mg_2SiO_4): Native as *olivine* or *chrysolite*. $\text{Mg}_3\text{Si}_2\text{O}_7 + 2\text{H}_2\text{O}$ found as *serpentine*, *talc*, *steatite* and *meerschaum*. Combined with Ca_2SiO_4 as *angite*, *hornblende*, *asbestos*, etc. Artificial as gelatinous $3\text{MgSiO}_3 + 5\text{H}_2\text{O}$.

Magnesium sulphide (MgS): Made by heating Mg and S to red heat. Decomposed by H_2O into $\text{Mg}(\text{OH})_2$ and $\text{Mg}(\text{SH})_2$; the latter known only in solution.

Magnesium nitride (Mg_3N_2): Greenish and amorphous. Decomposed by H_2O .

Phosphormagnesium (Mg_3P_2): Dark-brown and crystalline. Decomposed by H_2O .

Arsenicmagnesium (Mg_3As_2): Chocolate-brown; very brittle; easily decomposed.

Silicummagnesium (Mg_2Si): Lead-colored octohedra. With H_2O forms SiH_4 ; with HCl forms SiH_4 and SiO_2 .

Tests for magnesium salts:

1. Precipitated by sodium phosphate + ammonia as $\text{Mg}(\text{NH}_4)\text{PO}_4$.
2. Solubility of $\text{Mg}(\text{OH})_2$ or MgCO_3 in NH_4Cl separates Mg from the alkaline earths.

LVI. ZINC.

Symbol, Zn. Diad. Atomic weight, 64.905. Specific gravity, 7 to 7.2. Melts, 412° C. Boils, 940° C.

Occurs in ores as *blende* (ZnS), *calamine* (ZnCO₃), and ZnSiO₃. Made by roasting ores to remove CO₂ or convert ZnS into ZnO; then heated with charcoal in clay muffle or retorts and distilled by *Silesian method* or the English *destillatio per decensem*. The first product is *zinc dust* (Zn + ZnO) mixed with Cd, As, etc. Only the middle part of the distillate is pure. Commercial Zn is seldom pure. Purified by redistillation.

Properties of zinc: Bluish lustre, crystalline, hexagonal pyramids (also regular in presence of Cu). Brittle at ordinary temperatures; malleable and ductile between 100° and 150° C.; brittle above 200° C., and may be powdered. Oxidizes superficially in moist air. Burns in air or O with bluish light. Dissolves easily in dilute acids (best when impure). Liberates H. Reduces HNO to NH₃.

Zinc chloride (ZnCl₂): Anhydrous, formed by heating Zn in Cl. White; deliquescent; melts easily; boils at 676° to 683° C.; strong caustic and escharotic; easily soluble in water or alcohol. ZnCl₂ + H₂O forms from acid solutions. Decomposes in evaporating aqueous solutions. Used for baths of definite heat. Also in volumetric solution. Compound with K as ZnCl₂·2KCl.

Zinc bromide (ZnBr₂): Deliquescent.

Zinc iodide (ZnI₂): Deliquescent octohedra.

Zinc fluoride (ZnF₂): Rhombic crystals; little soluble.

Zinc silicofluoride (ZnSiF₆ + H₂O): Very soluble hexagonal crystals.

Zinc cyanide [Zn(CN)₂]: White insoluble powder. Forms double salts with alkaline cyanides.

Zinc oxide (ZnO): Native or *red zinc ore*. Zn burns with O to white oxide (*flores zinci*; *lava philosophica*), hexagonal crystals, yellow when hot; not volatile; insoluble in H₂O; specific gravity is 5.47; used as white paint.

Zinc hydrate [Zn(OH)₂]: Precipitated by KOH, etc., from Zn salts. Insoluble in H₂O; soluble in KOH, NaOH or NH₄OH; precipitated again by boiling; rhombic crystals.

Zinc nitrate [Zn(NO₃)₂ + 6H₂O]: Deliquescent quadratic crystals; soluble in alcohol or water.

Zinc carbonate (ZnCO₃): Native as *zinc spar* (*calamine*) in anhydrous hexagonal prisms. Also as amorphous *basic carbonate*. [Zinc blüthe, ZnCO₃ + 2Zn(OH)₂ + H₂O]; insoluble in H₂O and K₂CO₃ or Na₂CO₃; soluble in (NH₄)₂CO₃.

Zinc sulphate (white vitriol, ZnSO₄ + 7H₂O): Rhombic crystals, similar to MgSO₄. Monoclinic with 6H₂O. Crystal water differing (1 and 6) as with Mg salt: ZnSO₄·K₂SO₄ + 6H₂O.

Zinc sulphite ($\text{ZnSO}_3 + 2\text{H}_2\text{O}$): Little soluble in H_2O ; soluble in H_2SO_3 .

Zinc hyposulphite ($\text{ZnS}_2\text{O}_3 + 6\text{H}_2\text{O}$): Very soluble crystals.

Zinc phosphates:

The *primary* [$\text{ZnH}_4(\text{PO}_4)_2 + \text{H}_2\text{O}$]: Large triclinic crystals.

The *secondary* [$\text{ZnHPO}_4 + \text{H}_2\text{O}$]: Quite soluble.

The *tertiary* [$\text{Zn}_3(\text{PO}_4)_2 + 4\text{H}_2\text{O}$]: Made by adding Na_2HPO_4 to Zn salts; almost insoluble in H_2O ; soluble in acids.

Zinc arsenate: Similar to phosphate. *Adamine* is a basic salt in yellow or violet crystals from Chili ($\text{ZnAsO}_4 \cdot \text{ZnOH}$).

Zinc silicate (ZnSiO_4): As *witherite*; native hexagonal crystals. Native rhombic zinc glass ($\text{Zn}_2\text{SiO}_4 + \text{H}_2\text{O}$).

Zinc sulphide (ZnS): In regular crystals native as *zinc blende*; in hexagonal crystals native as *wurtzite*.

Zinc sulphhydrate [$\text{Zn}(\text{SH})_2$]: Precipitated by H_2S from neutral or alkaline solutions. Insoluble in H_2O or acetic acid; decomposed by strong acids; white.

There also exist: seleniumzinc (ZnSe); telluriumzinc (ZnTe); phosphor zinc (Zn_3P_2); arsenic zinc (Zn_3As_2) and antimony zinc (Zn_3Sb_2).

Tests for zinc salts:

1. Metal evolves H with acids (HCl or H_2SO_4).
2. Metal burns to white oxide.
3. Sulphide is white and soluble in acids.
4. Precipitated as $\text{Zn}(\text{OH})_2$ by NaOH , KOH , or NH_4OH ; soluble in excess of NH_4OH .
5. ZnCO_3 is soluble in NH_4OH . White precipitate with potassium ferrocyanide.
7. Bright green color formed with cobalt nitrate before blow-pipe.

LVII. CADMIUM.

Symbol, Cd. Diad. Atomic weight, 111.77. Specific gravity, 8.6. Melts, 321°C . Boils, 763 to 772°C . Atom occupies 2 volumes (molecular weight, 111.77).

Occurs with Zn ores and as *greenockite* (CdS). Obtained like Zn by distillation; the more volatile Cd comes over first. Discovered simultaneously by Stromeyer and Herman (1818). A white metal; forms octohedra; very ductile and malleable; vapor density at 940°C . = 3.94 (air = 1); burns in air to brown oxide; evolves H from HCl or H_2SO_4 ; forms fusible alloys (page 279, volume IX).

Cadmium chloride ($\text{CdCl}_2 + 2\text{H}_2\text{O}$): Formed by dissolving CdO in HCl . Anhydrous by efflorescence or heating. Sublimes in crystal scales. Melts at 541°C . Forms double salts.

Cadmium bromide ($\text{CdBr}_2 + 4\text{H}_2\text{O}$): Long anhydrous needles. Sublimes in scales.

Cadmium iodide (CdI_2): Hexagonal plates of nacreous luster. Melts at 404° and boils at 708° to 719° C. Forms double salts. Used in photography.

Cadmium fluoride (CdF_2): Little soluble.

Cadmium silicofluoride (CdSiF_6): Easily soluble.†

Cadmium cyanide [$\text{Cd}(\text{CN})_2$]: White. Insoluble in H_2O , but soluble in KCN, forming $\text{Cd}(\text{CN})_2 + 2\text{KCN}$ (colorless octohedra). Precipitated as CdS by H_2S .

Cadmium oxide (CdO): Brown amorphous, or by heating nitrate—octohedra. Blue-black by reflected, and brown by transmitted light. Insoluble in H_2O .

Cadmium hydrate [$\text{Cd}(\text{OH})_2$]: Formed by precipitating Cd salts by KOH. Insoluble in H_2O , KOH, or NaOH, but soluble in NH_4OH .

Cadmium nitrate [$\text{Cd}(\text{NO}_3)_2 + 4\text{H}_2\text{O}$]: Deliquescent prisms. Soluble in alcohol.

Cadmium carbonate (CdCO_3): White. Insoluble in $(\text{NH}_4)_2\text{CO}_3$.

Cadmium sulphate ($3\text{CdSO}_4 + 8\text{H}_2\text{O}$): Monoclinic. Has various other proportions of H_2O ; also double salts.

Cadmium phosphate [$\text{Cd}_3(\text{PO}_4)_2$]: Insoluble.

Cadmium arsenate [$\text{Cd}_3(\text{AsO}_4)_2$]: Insoluble.

Cadmium sulphide (CdS): Occurs native as yellow hexagonal *greenockite*. Precipitated by H_2S from acid solution as a yellow powder. Used as paint. Very little soluble in $(\text{NH}_4)_2\text{S}$.

Tests for cadmium salts:

1. Sulphide is yellow.
2. Oxide forms brown incrustations before blow-pipe.
3. Carbonate is insoluble in NH_4OH .
4. Precipitated as $\text{Cd}(\text{OH})_2$ by NaOH or KOH; insoluble in excess.
5. Precipitated as $\text{Cd}(\text{OH})_2$ by NH_4OH ; soluble in excess.
6. Precipitated by Zn.

LVIII.—BERYLLIUM.

Symbol, Be. Diad. Atomic weight, 9.085. Specific gravity, 1.64 to 2.1. Synonym, *glucinum*.

Discovered by Wöhler and Bussy (1828). Occurs in *beryll* [emerald, $\text{Al}^{ii}_2\text{Be}^{ii}_3(\text{SiO}_3)_6$], *chrysoberyll* (cymophane, $\text{Be}^{ii}\text{Al}^{iii}_2\text{O}_4$), *euclase*, *leucophane*, *helvite*, etc. Has a steel gray luster; unchanged in air; does not unite with O or S at red heat; liberates H from acids; does not decompose H_2O at 100° C.; forms a transition between Mg and Al.

Beryllium chloride (BeCl_2): colorless needles. $\text{BeCl}_2 + \text{H}_2\text{O}$ is deliquescent.

Beryllium bromide (BeBr_2), *beryllium iodide* (BeI_2), and *beryllium fluoride* (BeF_2) also exist.

Beryllium oxide (BeO): White insoluble powder.

Beryllium hydrate [$\text{Be}(\text{OH})_2$]: White insoluble powder.

Beryllium nitrate [$\text{Be}(\text{NO}_3)_2 \cdot 2\text{H}_2\text{O}$?]: Deliquescent.

Beryllium sulphate ($\text{BeSO}_4 + 4\text{H}_2\text{O}$): Quadratic octohedra.

Beryllium silicate (Be_2SiO_4): Hexagonal as *phenakit*.

Beryllium aluminium silicate [$\text{Be}^{11}_3\text{Al}^{111}_2(\text{SiO}_3)_6$]: Native as *beryll*, and when colored green by Cr as *emerald*.

Beryllium sulphide (BeS): Gray mass.

Be salts are precipitated by $(\text{NH}_4)_2\text{S}$ as BeO .

LIX. METALS OF THE ALUMINIUM GROUP.

	Al.	Ga.	In.	Tl.
Atomic weight.....	27.003	68.834	113.398	203.715
Specific gravity.....	2.670	5.900	7.300	11.800
Melting point.....	Red heat.	30° C.	176° C.	285° C.

They are all triads, or tetrads with double atoms linked by one atomicity, so as to leave three free atomicities.

LX. ALUMINIUM.

Symbol, Al. Tetrad (or triad). Atomic weight, 27.003. Specific gravity, 2.67. Melts, red heat. *Occurs* only in compounds, forming from 6 to 10 per cent of the earth's crust. Found native as oxide (ruby, sapphire, corundum, emery); hydrate (draspore, hydrargyllite, bauxite); silicate (clay, cyanite, feldspar); fluoride (cryolite) or phosphate (wavellite). *Made* largely from aluminium sodium chloride by heating with Na, or from cryolite with Na.

Properties of Aluminium: White luster; very malleable and ductile; not volatile; not oxidized in dry air; burns in O when used as thin foil; dissolves in HCl, liberating H; not affected by HNO_3 or dilute H_2SO_4 ; dissolves in KOH or NaOH with liberation of H.

Aluminium chloride [Al_2Cl_6 (or AlCl_3)]: Obtained by heating Al in Cl. Can not be obtained from aqueous solution. White hygroscopic hexagonal tables. Easily sublimed. Heat decomposes the deliquescent crystals of $\text{Al}_2\text{Cl}_6 + 12\text{H}_2\text{O}$ into Al_2O_3 and HCl. *Double chlorides* form.

Aluminium bromide [Al_2Br_6 (or AlBr_3)]: Similar to chloride.

Aluminium iodide [Al_2I_6 (or AlI_3)]: Similar to chloride.

Aluminium fluoride [Al_2F_6 (or AlF_3)]: Insoluble rhombohedra (nearly cubical). Greenland cryolite is $\text{AlF}_3 + 3\text{NaF}$.

Aluminium oxide (Al_2O_3): Native in very hard hexagonal crystals, as corundum, ruby, and sapphire. Emery is less pure, but very hard; almost insoluble in acids; must be fluxed with KHSO_4 or KOH.

Aluminium hydrate $[\text{Al}_2(\text{OH})_6]$: Native as hydrargyllite. Precipitated by NH_4OH from alum. Soluble in KOH or NaOH . May be obtained dialyzed in solution like Fe . Unites with organic colors to form lakes. With strong bases it unites, acting the part of an acid, and forms *aluminates*, which are derived from $\text{Al}_2\text{O}_2(\text{OH})_2$; native as diaspas, impure as bauxite; $\text{Al}_2\text{O}_2(\text{OK})_2$; $\text{Al}_2\text{O}_2(\text{ONa})_2$; $\text{Al}_2(\text{ONa})_6$; insoluble in H_2O ; decomposed by CO_2 . Aluminates of Ca , Sr , and Ba are little soluble. Aluminates of Mg group are native as spirrrels, pleonaste, sapphirine, etc.

Aluminium nitrate $[\text{Al}_2(\text{NO}_3)_6 + 18\text{H}_2\text{O}]$: Deliquescent monoclinic prisms.

Aluminium sulphate $[\text{Al}_2(\text{SO}_4)_3 + 18\text{H}_2\text{O}]$: Native as hairsalt, colorless, nacreous scales, very soluble, easily fused. *Aluminit* is a native basic salt; $\text{Al}_2\text{O}_2\text{SO}_4 + 9\text{H}_2\text{O}$. *Alums* are compounds with the alkaline sulphates. Native *alumite* $[\text{Al}_2\text{K}_2\text{O}_2(\text{SO}_4)_2 + 2\text{Al}_2\text{O}_2\text{SO}_4 + 6\text{H}_2\text{O}]$ is found in Hungary.

Potassium alum $[\text{Al}_2\text{K}_2(\text{SO}_4)_4 + 24\text{H}_2\text{O}]$: Octohedral crystals (or cubes when crystallized from alkaline solution) soluble in 7 parts H_2O at 15°C ., or $\frac{1}{2}$ part at 100°C . Alums of same crystalline form are found in which K is replaced by Na , NH_4 , Rb , Cs , or Tl ; Al may be replaced by Fe , Mn , Cr , Ga , or In ; Se may replace S .

Aluminium phosphate $[\text{Al}_2(\text{PO}_4)_2 + 8\text{H}_2\text{O}]$: Native as *neutral gibbsite*. A number of *basic phosphates* are native; rhombic *wavellite* $[2\text{Al}_2(\text{OH})_4\text{Al}_2\text{F}_2(\text{OH})_2 + 7\text{H}_2\text{O}]$; amorphous *kapnicite*; amorphous *turquoise* [halaite, $2\text{Al}_2\text{O}_3\cdot\text{P}_2\text{O}_5 + 5\text{H}_2\text{O}$] is blue with Cu .

Aluminium silicate (AlKSiO_8) : Occurs native in great variety. Anhydrous as *cyanite*, *andalusite*, *sillimanite*, *melanite* $(\text{Ca}_3\text{Al}_2\text{Si}_3\text{O}_{12})$, *topaz* $(5\text{Al}_2\text{SiO}_5 + \text{Al}_2\text{SiF}_{10})$; hydrated as *kaolin*, impure *clay*, *feldspar* $[\text{Al}_2\text{K}_2(\text{SiO}_4)_2 + 4\text{SiO}_2]$, and many other varieties]. Many double silicates. *Ultramarine* is a silicate of Al with S .

Aluminium alloys: Form with Mn , Fe , Ni , etc. Bronze is an alloy with Cu , similar in color to Au .

Tests for aluminium salts:

1. Precipitated by KOH or NaOH , is soluble in excess, but insoluble in carbonates or NH_4OH .

2. Blue mass when heated with cobaltous nitrate.

3. Precipitated as $\text{Al}_2(\text{OH})_6$ by $(\text{NH}_4)_2\text{S}$.

Porcelain: Made from kaolin and glazed with feldspar. *Earthenware* is glazed with NaCl .

LXI. GALLIUM.

Symbol, Ga . Tetrad (or triad). Atomic weight, 68.854. Specific gravity, 5.9. Melts, 80.1°C .

Occurs in zinc blende (about 0.002 per cent). Discovered by Lecogde

de Boisbaudran (1875), in blende from *Pierrefitte* (Pyrenees). A white, hard, little malleable metal. Melts at 30.1°C ., but does not solidify above 0°C . unless touched by a solid metal. Oxidized at red heat; does not affect boiling H_2O ; soluble in HCl , KOH , or NH_4OH , liberating H . Spectrum has two violet lines between G and H on both sides of hydrogen.

Gallium chlorides (GaCl_2 and GaCl_3): Crystals.

Gallium oxide (Ga_2O_3): Made by heating $\text{Ga}(\text{NO}_3)_3$. White and pliable.

Gallium hydrate [$\text{Ga}(\text{OH})_3$]: White, flocculent. Soluble in KOH or NH_4OH .

Gallium alum [$\text{GaNH}_4(\text{SO}_4)_2 + 12\text{H}_2\text{O}$ or $\text{Ga}_2(\text{NH}_4)_2(\text{SO}_4)_2 + 24\text{H}_2\text{O}$].

LXII. INDIUM.

Symbol, In. Tetrad (or triad). Atomic weight, 113.398. Specific gravity, 7.3. Melts, 176°C .

Occurs in some zinc blende (0.1 per cent). Reduced by ignition with H . White, lustrous, soft, and ductile. Not changed in air. Burns with violet flame and brown smoke to yellow In_2O_3 . Soluble in HCl , H_2SO_4 , or HNO_3 , liberating H .

Indium chloride (In_2Cl_6): White and deliquescent. Sublimes in crystals. Vapor density is 7.6 (air = 1).

Indium oxide (In_2O_3): Yellow. Brown while hot. Not fusible.

Indium hydrate [$\text{In}_2(\text{OH})_6$]: White and gelatinous. Soluble in KOH .

Indium nitrate [$\text{In}_2(\text{NO}_3)_6 + 8\text{H}_2\text{O}$]: Deliquescent.

Indium carbonate [$\text{In}_2(\text{CO}_3)_3$]: White. Soluble in $(\text{NH}_4)_2\text{CO}_3$.

Indium sulphate [$\text{In}_2(\text{SO}_4)_3$]: Not crystallizable. Forms double salts.

Indium alum [$\text{In}_2(\text{NH}_4)_2(\text{SO}_4)_4 + 24\text{H}_2\text{O}$].

Indium was discovered by Reich & Richter (1863), by its spectrum of one blue and one violet line.

LXIII. THALLIUM.

Symbol, Tl. Triad (and monad), atomic weight, 203.715. Specific gravity, 11.8 to 11.9. Melts, 285°C . Discovered by Crookes and by Lamy (1861). *Occurs* in pyrites and in water. Forms triad and monad compounds. The latter resemble the alkalies. Is a white, soft metal. Oxidizes in air; preserved under H_2O .

Thallic chloride (TlCl_3) and *thallous chloride* (TlCl).

Thallic bromide (TlBr_3) and *thallous bromide* (TlBr).

Thallic iodide (TlI_3) and *thallous iodide* (TlI): The latter is an insoluble yellow powder. Colors green in sunlight. (Iodine reagent).

Thallic oxide (Tl_2O_3): By burning in O. Black.

Thallous oxide (Tl_2O): Formed by oxidization in air. Black.

Thallic hydrate [$Tl(OH)_3$]: Brown insoluble powder.

Thallous hydrate [$TlOH + 2H_2O$]: Colorless; easily soluble; alkaline reaction.

Thallic nitrate [$Tl(NO_3)_3 + 8H_2O$]: Easily decomposed by much H_2O .

Thallous nitrate [$TlNO_3$]: Rhombic prisms. Soluble in 10 parts of H_2O .

Thallic sulphate [$Tl_2(SO_4)_3 + 7H_2O$] and *thallous sulphate* (Tl_2SO_4): There are neutral, acid, and double salts.

Thallic sulphide (Tl_2S_3): Formed only in dry way. Black and fusible. Double sulphides form (as TlS_2K).

Thallous sulphide (Tl_2S): Precipitated by H_2S or alkaline sulphides. Black; amorphous; insoluble in H_2O ; soluble in HNO_3 or H_2SO_4 .

Thallous carbonate (Tl_2CO_3): Colorless and monoclinic. Both acid and neutral salts form.

Thallous chromate (Tl_2CrO_4) and *thallous dichromate* ($Tl_2Cr_2O_7$).

Thallium flint-glass is used for very high dispersion.

LXIV.—IRON GROUP.

	Cr.	Mn.	Fe.	Ni.	Co.
Atomic weight.....	52.009	53.906	55.913	57.928	58.887
Specific gravity.....	6.810	7.140	7.840	8.800	8.500

From a consideration of the vapor densities of $FeCl_6$ (11.39) and Al_2Cl_6 (9.34), the metals of the Fe group and Al are supposed to be *tetrads* linked by one atomicity in pairs ($\equiv Fe^{IV}-Fe^{IV} \equiv$); others consider from $Al(CH_3)_3$ that they are *triads*. They form diatomic compounds as well as triatomic; $=Fe^{IV}=Fe^{IV}=$ or $=Fe^{III}-Fe^{III}=$. The diatomic salts are most pronounced in Ni and least in Cr, in the following order: diatomic Ni, Co, Mn, Fe, Cr triatomic.

LXV. NICKEL.

Symbol, Ni. Tetrad (or triad). Atomic weight, 57.928. Specific gravity, 8.8 to 9.1. Melts, $1,600^\circ C$. Occurs in metallic state in meteorites; with As in Kupfer nickel, *arsenic nickel* (hence name); with S and As in *nickel glance*; with S in *hair pyrites*; with S and Sb in *nickel antimony glance*; also as carbonate, silicate, etc. Generally accompanied by Co. Sold in cubes (impure metal).

Properties of nickel: Silvery white with faint yellow luster; very hard; high polish; brittle; becomes malleable by adding 0.3 per cent of P or $\frac{1}{16}$ per cent of Mg; hard to fuse; magnetic; unchanged in air

or water, hence used for plating; slowly soluble in HCl or dilute H_2SO_4 ; readily soluble in HNO_3 .

Nickel alloys: Argentan (German silver); Cu 8, Zn $3\frac{1}{2}$, and Ni 2, 3 or 4 parts. Coin—Cu $\frac{3}{4}$ and Ni $\frac{1}{2}$ part, or (billon) Ni, Cu, Zn, and Ag.

Nickel is found with Fe in meteoric iron.

The oxygen salts are only nickelous (Ni^{II}).

Nickel chloride (NiCl_2): Anhydrous is yellow, but $\text{NiCl}_2 + 6\text{H}_2\text{O}$ is green. Monoclinic. Soluble in H_2O or alcohol. Forms double salts.

Nickel bromide (NiBr_2): Similar to chloride.

Nickel iodide (NiI_2): Similar to chloride.

Nickel cyanide [$\text{Ni}(\text{CN})_2$]: Greenish-white; insoluble in H_2O ; soluble in KCN forming *double cyanides* with Na, Ba, or K, from which dilute HCl precipitates $\text{Ni}(\text{CN})_2$.

Nickelous oxide (NiO): Occurs native as brown regular crystals. Made by heating hydrate.

Nickelic oxide (Ni_2O_3): Made by heating $\text{Ni}(\text{NO}_3)_2$.

Ni_3O_4 is a gray crystalline powder.

Nickelous hydrate [$\text{Ni}(\text{OH})_2$]: Green. Soluble in NH_4OH or acids.

Nickelic hydrate [$\text{Ni}(\text{OH})_3$]: Made by passing Cl into H_2O containing $\text{Ni}(\text{OH})_2$. Brownish-black. Amorphous.

Nickel nitrate [$\text{Ni}(\text{NO}_3)_2 + 5\text{H}_2\text{O}$]: Green, monoclinic, and deliquescent. Double salts with NH_4 are blue octohedra.

Nickel nitrite [$\text{Ni}(\text{NO}_2)_2$]: Orange green in solutions. Double salts as $\text{NiBaK}_2(\text{NO}_2)_6$; NiSrK_2 ; NiCaK_2 .

Nickel carbonate ($\text{NiCO}_3 + 6\text{H}_2\text{O}$): Green. Insoluble in H_2O or Na_2CO_3 . Soluble in NH_4OH (blue).

Nickel sulphate ($\text{NiSO}_4 + 7\text{H}_2\text{O}$): Dark-green rhombic crystals. When crystallized above 40°C . forms blue-green monoclinic, $\text{NiSO}_4 + 6\text{H}_2\text{O}$. Above 50°C . forms quadratic crystals. Very soluble. Forms double salts with alkaline sulphates, used in nickel-plating.

Nickel phosphate [$\text{Ni}_3(\text{PO}_4)_2 + x\text{H}_2\text{O}$]: Green.

Nickel arsenate [$\text{Ni}_3(\text{AsO}_4)_2 + 8\text{H}_2\text{O}$]: Native as *nickel bloom*.

Nickel silicate: Occurs native as *garnierite* in New Caledonia.

Nickel sulphide (NiS): Native as *hair pyrites* (millerite). Artificial is black. Slightly soluble in $(\text{NH}_4)_2\text{S}$.

Tests for nickel salts:

1. Not precipitated from acid solution by H_2S .
2. Precipitated black by $(\text{NH}_4)_2\text{S}$.
3. Hydrrous salts are green and anhydrous yellow.
4. Forms yellow borax bead.

LXVI. COBALT.

Symbol, Co. Tetrad (or triad). Atomic weight, 58.887. Specific gravity, 8.5 to 8.9. Occurs usually with Ni. White with reddish luster;

very ductile; polishable; magnetic; less fusible than Ni; unaltered by air or H_2O ; soluble like Ni in acids. The salts are similar to Ni salts, but red when hydrous and blue when anhydrous.

Cobalt chloride ($CoCl_2$ and $CoCl_2 + 6H_2O$): Used as sympathetic ink (cyanide still better). Also double salts.

Cobaltous cyanide [$Co(CN)_2 + xH_2O$]: Brown.

Cobaltic cyanide: Only in compounds.

Cobaltous oxide (CoO): Olive green. Soluble in acids.

Cobaltic oxide (Co_2O_3): Brownish-black.

Co_3O_4 is black.

Cobaltous hydrate [$Co(OH)_2$] and *cobaltic hydrate* [$Co_2(OH)_6$].

Cobaltous nitrate [$Co(NO_3)_2 + 6H_2O$]: Deliquescent monoclinic crystals. Much used as blow-pipe reagent.

Cobaltous nitrite [$Co(NO_2)_2$]: Very unstable. Double salt with K.

Cobaltic nitrite: Only in double salts.

Potassio-cobaltic nitrite [Fischer's salt, $Co_2(NO_2)_6 + 6KNO_2 + xH_2O$]: Yellow. Very slightly soluble in warm H_2O . Insoluble in potassium nitrite or acetate.

Sodio-cobaltic nitrite [$Co_2(NO_2)_6 + 6NaNO_2$]: Reagent for K; soluble.

Cobaltous carbonate ($CoCO_3$): Anhydrous as *cobalt spar*. With $6H_2O$ is in prisms.

Cobaltous sulphate ($CoSO_4 + 7H_2O$): Red and monoclinic. Also with less H_2O .

Cobalt phosphate [$Co_3(PO_4)_2$ and $CoHPO_4 + H_2O$].

Cobalt arsenate: Native as *cobalt bloom*.

Cobalt silicate: Made from *Zaffer* (roasted ore), quartz and K_2CO_3 . In blue glass, smaltz, etc.

Cobalt sulphide (CoS): Is similar to NiS .

Cobaltamines or *cobaltiac* salts are formed with NH_3 .

Dichro-cobalt chloride [dichroic, $Co_3(NH_3)Cl_3 + H_2O$]: Hexagonal.

Praseo-cobalt chloride [cobalt tetramine, $Co_4(NH_3)Cl_3 + H_2O$]: Green.

Purpureo-cobalt chloride [pentamine, $Co_5(NH_3)Cl_3$]: Purple; quad-ratic.

Roseo-cobalt chloride: Same as above with H_2O .

Luteo-cobalt chloride [hexamine, $Co_6(NH_3)Cl_3$]: Brownish-yellow.

Xantho-cobalt salts: $CoNO_2 \cdot 5(NH_3) \cdot (NO_3)_2$; $CoNO_2 \cdot 5(NH_3) \cdot SO_4$; also chloride.

Flavo-cobalt chloride [$CoNO_3 \cdot NO_2 \cdot (NH_3)_5Cl$].

Tests for cobalt salts:

1. Blue borax lead.
2. Blue Al compound before blow-pipe.
3. Green Zn compound before blow-pipe.
4. $BaCO_3$ precipitates Co from chlorinated solutions of Co and Ni, leaving Ni in solution. Co was discovered in 1733 by Brandt.

LXVII. MANGANESE.

Symbol, Mn. Tetrad (or triad). Atomic weight, 53.906. Specific gravity, 7.14 to 7.99.

Occurs in great abundance, but not in metallic state. Native oxides: *Pyrolusite*, *braunite*, *hausmannite*, *manganite*, *psilomelan*. Sulphide: *Manganese blende*. Carbonate: *Manganese spar*.

Discovered by Scheele (1774). Grayish metal with reddish tint, very hard and brittle. Melts at high temperatures. Rapidly oxidized in moist air, decomposes H_2O (when in powder) at ordinary temperatures. Dissolves in acids, liberating H. When containing Si it is very unchangeable and takes a high polish.

Manganous chloride ($MnCl_2$): A red crystalline, fusible mass. $MnCl_2 + 4H_2O$ is red, monoclinic. Double salts form.

Manganic chloride (Mn_2Cl_6): Not isolated and exists only in solution.

Manganous cyanide [$Mn(CN)_2$] and *manganic cyanide* [$Mn_2(CN)_6$]: Exist only in compounds.

Mangano-cyanide of potassium [$Mn(CN)_2 + 4KCN + 3H_2O$]: From solutions I precipitates $Mn_2(OH)_6$ (separation of Mn from Fe). A similar Na salt forms.

Mangani-cyanide of potassium [$Mn_2(CN)_6 + 6KCN$]: Red rhombic prisms. Easily soluble. Solutions readily decomposed.

Manganous oxide (monoxide, MnO): Native *manganosite* in regular green or red octohedra. Artificial is green. Easily oxidized unless ignited.

Manganic oxide (Mn_2O_3): Native *braunite* as quadratic octohedra. Colors glass violet.

Manganoso-manganic oxide (Mn_3O_4): Native *hausmannite* as black quadratic octohedra.

Manganese dioxide (MnO_2): Native as rhombic *pyrolusite*. On heating yields O. With HCl yields Cl.

Manganous hydrate [$Mn(OH)_2$]: Formed as pure white precipitate by KOH, but rapidly oxidizes, becoming brown.

Manganic hydrate [$Mn_2(OH)_6$]: Brownish-black.

$MnO(OH)$ is native as rhombic black *manganite*. MnO_3H_2 ; $Mn(OH)_4$; $Mn_3O_8H_4$ also exist.

$Mn_4O_{10}H_3K$ is brown and insoluble. Forms during oxidation of organic bodies by $KMnO_4$ in neutral solutions.

Manganous nitrate [$Mn(NO_3)_2 + 6H_2O$]: Very deliquescent and unstable.

Manganous carbonate ($MnCO_3$): Formed as a white precipitate by alkalis.

Manganous sulphate ($MnSO_4 + 7H_2O$): Peach-bloom colored monoclinic crystals when formed below $6^\circ C$. At 15° has $5H_2O$ (like $CuSO_4$); between 20° to $30^\circ C$. has $4H_2O$.

$\text{MnSO}_4\text{K}_2\text{SO}_4 + 6\text{H}_2\text{O}$ is monoclinic.

Manganese phosphate $[\text{Mn}_3(\text{PO}_4)_2 + 7\text{H}_2\text{O}]$: Red rhombic prisms. Also *acid salts* and *double salts*.

Manganous silicate $(\text{Mn}_2\text{SiO}_4)$: Native as *theophoite*. Also MnSiO_3 (rhodonite), etc. Colors the amethyst.

Manganic sulphate $[\text{Mn}_2(\text{SO}_4)_3]$: A green powder.

Manganic acid (H_2MnO_4) : Mineral chameleon; green-black; rhombic; easily soluble; oxidized to permanganate.

Barium manganate (BaMnO_4) .

Permanganic acid (HMnO_4) : Only the anhydrous (Mn_2O_7) known in the free state. Very unstable. Detonates on heating. Ignites organic compounds.

Potassium permanganate (KMnO_4) : Deep purple rhombic prisms. Soluble in 16 parts H_2O at 15°C . (solution has deep violet color). Strong oxidizer. Used in volumetric analysis.

NaMnO_4 ; NH_4MnO_4 ; $\text{Ba}(\text{MnO}_4)_2$; AgMnO_4 also exist.

Manganous sulphide (MnS) : Natural black cubes as *manganese blende*. Artificial is a green powder. The flesh-colored hydrated $(\text{MnS} + \text{H}_2\text{O})$ is precipitated by alkaline sulphides. Easily oxidized.

Manganic sulphide (MnS_2) : Native as red brown *hanerite* (regular octohedra).

Manganese alloy with Fe (spiegeleisen) is used in the Bessemer steel process.

Tests for manganese salts:

1. Sodium manganate is blue before the blow-pipe.
2. Sulphide is flesh-colored.

LXVIII. IRON.

Symbol, Fe. Tetrad (or triad). Atomic weight, 55.913. Specific gravity, 7.84. Melts, 1800°C .

Occurs in metallic state in meteorites. Also found in plants and blood of animals. The *ores* are oxides, hydrates, carbonates, and sulphides. Made from ores by melting with carbon to reduce, and addition of lime and sand to dissolve the impurities.

Cast-iron: Contains 3 to 6 per cent of C. The white, hard, and brittle grade is crystalline. The gray is more granular.

Wrought iron: Contains 0.2 to 0.6 per cent of C. Is decarbonized in the process of manufacture by puddling, etc. Has high fusion point. Can be welded.

Malleable cast-iron: Similar in properties to wrought iron.

Steel: Contains 0.9 to 2 per cent of C. Is elastic. Is annealed and hardened.

Pure iron (hydrogenio redactum): Specific gravity, 7.75 to 7.78. Silver white. Polishable. Softer than wrought iron. Has very high fusing point. Does not change in dry air, but rusts in moist air with CO_2 . Burns in air to *ferroso-ferric* or *magnetic oxide* (Fe_3O_4). In fine powder decomposes H_2O at 15°C .; more readily at 100° , and still more so at red heat. Dissolves in dilute HCl or H_2SO_4 liberating H . Concentrated HNO_3 dissolves it.

Ferrous chloride (FeCl_2): White and fusible. The hydrous [$\text{FeCl}_2 + 4\text{H}_2\text{O}$ (or $2\text{H}_2\text{O}$)] is green. Double alkaline salts form as $\text{FeCl}_2 + 2\text{KCl} + 2\text{H}_2\text{O}$.

Ferric chloride (Fe_2Cl_6): Made by heating Fe in Cl . Is brown, deliquescent, hexagonal, and fusible. Vapor density between 440° to 600°C . is 11.39 to 11.01. Soluble in alcohol or ether. The hydrous is $\text{Fe}_2\text{Cl}_6 + 12\text{H}_2\text{O}$. *Tinctura ferri chloridi* contains 13.23 per cent, and *liquor ferri chloridi* 37.8 per cent Fe_2Cl_6 . When aqueous solution is heated, the salt decomposes into soluble colloidal ferric hydrate [$\text{Fe}_2(\text{OH})_6$] and HCl , but reunites on cooling. Deposits oxychloride on standing. May be dialysed to separate HCl .

Ferrous bromide (FeBr_2) and **ferric bromide** (Fe_2Br_6): Are similar to the Cl compounds.

Ferrous iodide (FeI_2): Made by direct union of Fe and I . Gray. The hydrous ($\text{FeI}_2 + 4\text{H}_2\text{O}$) is pale, green crystals. *Syrupus ferri iodidi* contains 10 per cent FeI_2 .

Ferric iodide (Fe_2I_6): Not known in isolated state

Ferrous fluoride ($\text{FeF}_2 + 8\text{H}_2\text{O}$): Pale green. Anhydrous.

Ferric silicofluoride (FeSiF_6): Bluish-green and very soluble.

Ferric fluoride (Fe_2F_6): Colorless. Hydrous has $9\text{H}_2\text{O}$.

Ferrous cyanide [$\text{Fe}(\text{CN})_2$] and **ferric cyanide** [$\text{Fe}_2(\text{CN})_6$]: Only known in compounds.

Potassium ferrocyanide [$\text{K}_4\text{Fe}^{II}(\text{CN})_6 + 3\text{H}_2\text{O}$]: Made from leather, etc. Yellow quadratic crystals. Soluble in 3 to 4 parts H_2O . Heated to redness it forms KCN , N , and Fe_2C . Heated with dilute H_2SO_4 it yields HCN and $\text{KFe}(\text{CN})_5$. Forms white precipitate with ferrous salts, turning blue in air. Forms prussian blue, $3[\text{Fe}^{II}(\text{CN})_2] + 4[\text{Fe}^{III}(\text{CN})_3]$, with ferric salts. The precipitate with Cu is red-brown; U , brown; Ni , pale-blue; Co , green; Ag , Cd , Zn , Mn , Pb , Hg , white.

Potassium ferricyanide [$\text{K}_3\text{Fe}^{III}(\text{CN})_6$]: Formed from ferrocyanide solutions by action of Cl . Crystals are red and monoclinic. Soluble. Reduced by solar light. Ferrous salts precipitate Turnbull's blue, $3[\text{Fe}^{II}(\text{CN})_2] + \text{Fe}_2^{III}(\text{CN})_6$. Ferric salts form brown solutions.

Sodium nitroprusside [$\text{FeNa}_2(\text{NO})(\text{CN})_5 + 2\text{H}_2\text{O}$]: Gives purple color with soluble sulphides.

Ferrous oxide (FeO): Black powder. Easily oxidized. Reduced by H .

Ferric oxide (Fe_2O_3): Native as *iron glance* (specular) or *red hæmatite*. Made by heating native or artificial hydroxide. Very little soluble in acids.

Ferroso-ferric oxide (Fe_3O_4): Native as magnetic ore in octohedra. Smithy scales are similar.

Ferrous hydrate [$\text{Fe}(\text{OH})_2$]: Formed from ferrous salts as a white precipitate if free from air. Very unstable. Oxidizes, becoming green, then black, and finally red.

Ferric hydrate [$\text{Fe}_2(\text{OH})_6$ or $\text{Fe}(\text{OH})_3$]: Formed as a reddish-brown precipitate from solutions of Fe_2Cl_6 by NH_4OH . Becomes brown and denser by boiling; forming $\text{Fe}_2\text{O}(\text{OH})_4$, then $\text{Fe}_2\text{O}_2(\text{OH})_2$, and lastly Fe_2O_3 . Is soluble when freshly made from solution of Fe_2Cl_6 or ferric acetate, from which it may be obtained in aqueous solution by dialysis. When freshly prepared it is used as an *antidote to arsenic*. $\text{Fe}_2\text{O}_2(\text{OH})_2$ acts as a feeble monobasic acid; *magnetic iron* ($\text{Fe}^{\text{iii}}\text{O}-\text{O}-\text{Fe}^{\text{ii}}\text{O}-\text{FeO}$) being one of its salts. *Franklinite* contains Zn and Mn in place of Fe^{ii} ; *delafossite* Cu as $\text{Fe}^{\text{iii}}\text{O}-\text{O}-\text{Cu}$. There is also $\text{Fe}^{\text{iii}}=\text{O}-\text{O}-\text{CaO}-\text{Fe}=\text{O}$. The precipitates from solutions of Fe_2Cl_6 by KOH or NaOH contains similar salts; $\text{Fe}_2^{\text{iii}}=\text{O}_2-(\text{OH})_2$.

Ferrous nitrate [$\text{Fe}(\text{NO}_3)_2 + 6\text{H}_2\text{O}$]: Unstable crystals. Rapidly oxidized in air.

Ferrous carbonate (FeCO_3): Native as *spathic iron ore* in colorless, yellow, or brown hexagonal forms. *Sphaerosiderit* contains Al. Dissolves in water containing CO_2 , forming *chalybeate waters*.

Ferrous sulphate ($\text{FeSO}_4 + 7\text{H}_2\text{O}$): Bluish-green monoclinic efflorescent crystals. Soluble in $1\frac{1}{2}$ parts H_2O at 15° or $\frac{1}{2}$ part at 100°C . Oxidizes in contact with H_2O . Double salts form as: $\text{FeSO}_4 \cdot \text{K}_2\text{SO}_4 + 6\text{H}_2\text{O}$.

Ferrous phosphate [$\text{Fe}_3(\text{PO}_4)_2 + 8\text{H}_2\text{O}$]: *Vivianite* is in green or blue scales, but the amorphous white is precipitated by Na_2HPO_4 , turning blue in the air.

Ferrous acid phosphate [$\text{FeH}_4(\text{PO}_4)_2$]: Rhombic.

Ferrous silicate (Fe_2SiO_4): Native as *fayalite*. Occurs in slags, etc.

All ferrous salts give blue reaction with ferricyanide and colorless with sulphocyanide.

Ferric nitrate [$\text{Fe}_2(\text{NO}_3)_6 + 12\text{H}_2\text{O}$]: Cubes; or with $9\text{H}_2\text{O}$, monoclinic prisms. Deliquescent. Solutions on boiling deposit brown insoluble basic salts. The *liquor ferri nitratis* contains 6 per cent of anhydrous $\text{Fe}_2(\text{NO}_3)_6$.

No ferric carbonate.

Ferric sulphate [$\text{Fe}_2(\text{SO}_4)_3 + 9\text{H}_2\text{O}$]: Hexagonal prisms. *Liquor ferri tersulphatis* contains 28.7 per cent of $\text{Fe}_2(\text{SO}_4)_3$, and *liquor ferri subsulphatis*, 43.7 per cent of $\text{Fe}_4\text{O}(\text{SO}_4)_5$. Double salts (alums) form with K, NH_4 , etc.

Ferric phosphate $[\text{Fe}_2(\text{PO}_4)_2 + 4\text{H}_2\text{O}]$: White, and soluble in acids (except acetic).

Ferric acid phosphate $[\text{Fe}^{\text{III}}\text{H}_6(\text{PO}_4)_3]$: Red deliquescent rhombic crystals.

Ferric silicate $[\text{FeH}(\text{SiO}_3)_2]$: Native as *anthosiderite*.

Ferric acid (H_2FeO_4) : Only known in salts.

Potassium ferrate (K_2FeO_4) : Small deep-red crystals, soluble in H_2O . Decomposed by heat.

Sodium ferrate $(\text{Na}_2\text{FeO}_4)$: Similar to K salt.

Barium ferrate (BaFeO_4) : Stable salt. Red. Insoluble in H_2O .

Ferrous sulphide (FeS) : Occurs as *troilite* in meteoric iron. Made by direct union or by H_2S in alkaline solution. Black. With acids, evolves H_2S .

Magnetic ferrous sulphide: This is FeS with a slight excess of S.

Ferric sulphide (Fe_2S_3) : A grayish-yellow powder.

Ferrum disulphide (FeS_2) . Native as *pyrites* (regular) or *marcasite* (rhombic). On heating yields S and FeS .

Arsenic-iron (FeAs_2) : Native.

Ferrum sulph-arsenite (FeAsS) : Native as *Misspickel* in rhombic crystals.

All *ferric salts* give red color with sulphocyanide and blue reaction with ferrocyanide.

LXIX. CHROMIUM.

Symbol, Cr. Tetrad (or triad). Atomic weight, 52.009. Specific gravity, 6.81.

Occurs in chromic iron ore $(\text{FeO} \cdot \text{Cr}_2\text{O}_3)$ and red lead ore, also as coloring in some minerals. Gray minute rhombohedra. Not magnetic. Oxidizes slowly at red heat and burns in O. Dissolves in cold HCl or hot H_2SO_4 but not in HNO_3 .

Chromous chloride (CrCl_2) : White. Dissolves in H_2O with blue color; absorbs O and becomes green.

Chromic chloride $(\text{Cr}_2\text{Cl}_6 \text{ or } \text{CrCl}_3)$: Peach-bloom colored scales made in dry way. When heated in air becomes Cr_2O_3 . Insoluble in H_2O when pure unless CrCl_2 is present. Also hydrated $(\text{Cr}_2\text{Cl}_6 + 12\text{H}_2\text{O})$.

Chromous bromide (CrBr_2) and *chromic bromide* (Cr_2Br_6) : Similar to chlorides.

Chromic fluoride (Cr_2F_6) .

No *chromium monoxide*.

Chromic oxide (Cr_2O_3) : Made by heating the hydrate. Dark-green, hard, hexagonal crystals. Almost insoluble after ignition. Colors glass and borax beads green.

Chromous hydrate [$\text{Cr}(\text{OH})_2$]: Forms as yellow precipitates from CrCl_2 by KOH .

Chromic hydrate [$\text{Cr}_2(\text{OH})_6$]: Forms as precipitate from chromic salts by NH_4OH . The hydrous [$\text{Cr}_2(\text{OH})_6 + 4\text{H}_2\text{O}$] is pale blue. KOH or NaOH forms green precipitates containing some K or Na .

$\text{Cr}_2\text{O}_3(\text{OH})_2$ is grayish-blue and insoluble in HCl .

$\text{Cr}_2\text{O}(\text{OH})_4$ is *Guignet's green*. Made by melting $\text{K}_2\text{Cr}_2\text{O}_7$ with 3 parts boric acid and dissolving out the borate.

$\text{Cr}_2\text{O}_3(\text{OH})_2$ acts as an acid in some compounds; chromic iron ore, $\text{CrO}-\text{OFeO}-\text{OCr}$. Chromium forms compounds with Ba , Mn , or Zn .

Chromous oxygen salts: Very unstable. Only one studied.

Chromous sulphate ($\text{CrSO}_4 + 7\text{H}_2\text{O}$): Blue monoclinic crystals. Absorbs O . $\text{CrSO}_4 + \text{K}_2\text{SO}_4 + 6\text{H}_2\text{O}$ is monoclinic.

Chromic nitrate [$\text{Cr}_2(\text{NO}_3)_6 + 18\text{H}_2\text{O}$]: Soluble red monoclinic crystals.

Chromic sulphate [$\text{Cr}_2(\text{SO}_4)_3 + 18\text{H}_2\text{O}$]: Solutions are first green and do not crystallize, but become gradually violet and crystallize in blue octohedra. Soluble in H_2O , but precipitated by alcohol. The aqueous solution on heating decomposes into green basic salt and free acid, and is not precipitated by alcohol until changed to violet again by standing.

Chromic alum [$\text{Cr}_2(\text{SO}_4)_3 + \text{K}_2\text{SO}_4 + 24\text{H}_2\text{O}$]: Deep purple, nearly black octohedra. Also ammonio chromic alum.

Chromic acid (H_2CrO_4): Not isolated.

Chromic anhydride (Cr_2O_6): Formed from $\text{K}_2\text{Cr}_2\text{O}_7$ and H_2SO_4 . Red rhombic, deliquescent, soluble crystals. Great oxidizer. Inflames alcohol. Gives blue reaction with H_2O_2 , which is supposed to contain perchromic acid, soluble in ether.

Potassium chromate (K_2CrO_4): Yellow rhombic crystals soluble in 2 parts of H_2O . Used as indicator in volumetric analysis.

Potassium dichromate ($\text{K}_2\text{Cr}_2\text{O}_7$): Made on large scale by melting chromic iron with KNO_3 , etc. Red triclinic crystals, soluble in 10 parts of H_2O at 15°C . When treated with HNO_3 *tri* and *tetra chromates* are formed.

Na and NH_4OH salts of Cr are similar to the K salts.

Barium chromate (BaCrO_4): Yellow. Insoluble in H_2O or acetic acid, but soluble in HCl .

Strontium chromate (SrCrO_4): Insoluble in H_2O , but soluble in acetic acid.

Calcium chromate (CaCrO_4): Soluble in H_2O .

Mg , Zn , Cd , Ni , Co , Cu , and Mn form basic chromium salts, insoluble in H_2O .

Silver chromate (Ag_2CrO_4): Red amorphous or dark-green crystals.

Mercuric chromate (HgCrO_4) and red *mercurous chromate* (Hg_2CrO_4). Also basic salts.

Chromium oxychloride (CrO_2Cl_2): Red volatile liquid. Decomposed by H_2O into HCl and H_2CrO_4 . Forms KClCrO_3 ; KBrCrO_3 ; KFCrO_3 ; etc.

Chromous sulphide (CrS): Black powder.

Chromic sulphide (Cr_2S_3): Formed direct from metallic Cr and S. Gray.

Chromium nitrogen (CrN): Made by heating Cr in N. Black powder.

Chrom-amine compounds: Analogous to cobalt-amines.

Tests for chromic acid salts:

1. Pale yellow precipitate with Ba salts.
2. Bright-yellow precipitate with Pb salts.
3. Brick-red precipitate with mercurious salts.
4. Crimson precipitate with Ag salts.
5. Green oxide by reductions.

LXX. MOLYBDENUM GROUP.

	Mo.	W.	Ur.
Atomic weight.....	95.888	183.610	238.482
Specific gravity.....	8.600	19.129	18.830

This group is somewhat related to the Cr group.

The members are all hexads.

LXXI. MOLYBDENUM.

Symbol, Mo. Hexad. Atomic weight, 95.888. Specific gravity, 8.6. A rare metal discovered by Scheele (1778). Occurs as *molybdenite* (MoS_2), *molybdenum ochre* (MoO_3) and *lead molybdate* (PbMoO_4). Is silver white. Less fusible than Pt. Very hard. Soluble in concentrated H_2SO_4 , HNO_3 , or nitrohydrochloric acid; when heated in air it oxidizes to MoO_3 .

Molybdenum dichloride (MoCl_2): Yellow insoluble powder.

Molybdenum trichloride (MoCl_3): Brownish-red amorphous powder.

Molybdenum tetrachloride (MoCl_4): Brown powder.

Molybdenum pentachloride (MoCl_5): Made by heating Mo in Cl_2 . With H_2O forms brown liquid becoming colorless by dilution. Solution in alcohol, ether, or HCl is green; in concentrated H_2SO_4 blue; in HNO_3 colorless.

Molybdenum also forms *oxychlorides*, *bromides*, *oxybromides*, and *fluorides*.

Molybdenum dioxide (MoO_2): Brown crystals.

Molybdenum trioxide (MoO_3): White (yellow while hot). Insoluble in H_2O or acids, but easily soluble in caustic alkalis.

Mo_2O_3 is black and insoluble in acids.

Molybdenum hydroxides: $\text{Mo}(\text{OH})_3$, black; $\text{Mo}(\text{OH})_4$, brown.

Molybdic acid $[\text{MoO}_2(\text{OH})_2 \text{ and } \text{Mo}_5\text{O}_{16}\text{H}_2 = \text{MoO}_2(\text{OH})_2 + 4\text{MoO}_3]$: Colorless solution in HCl , but reduction with Zn turns first blue, then green, and brown.

Ammonium molybdate $[(\text{NH}_4)_2\text{MoO}_4]$: Used in determination of phosphoric acid.

$\text{MoO}_2(\text{OH})_2$ also forms salts with K , Na , Mg , Zn , and Pb .

Molybdenum sulphate $(\text{MoO}_2\text{SO}_4)$: This is the only Mo salt with O acids studied.

Phospho-molybdic acid $(11\text{MoO}_3 + \text{PO}_4\text{H}_3?)$: Other formulæ suggested.

LXXII. TUNGSTEN.

Symbol, W . Hexad. Atomic weight, 183.61. Specific gravity, 19.129. Synonym, *wolframium*.

Occurs as ferrous tungstate in *wolfram*, combined with Ca in *scheelite*, with Pb as *lead tungstate*. The metal is obtained by reduction in H . Hard, brittle, little fusible, quadratic crystals. Burns in O to WO_3 . Used to alloy steel.

Tungsten chlorides: WCl_2 , WCl_4 , WCl_5 , and WCl_6 . Also *oxychlorides*; WOCl_4 and WO_2Cl_2 .

Tungsten also forms *bromides*, *iodides*, and *fluorides*.

Tungsten dioxide (WO_2) : Brown powder. Little affected by acids.

Tungsten trioxide (tungstic anhydride, WO_3): Yellow. Insoluble in H_2O or acids, but soluble in KOH .

Tungsten hydroxide $[\text{W}(\text{OH})_4]$: Blue. Reduced from tungstic acid by Zn .

Tungstic acid $[\text{WO}(\text{OH})_4 \text{ and } \text{WO}_2(\text{OH})_2]$: Reduced by H from Zn and HCl , first blue, then brown. Soluble. Forms *tungstates*.

Metatungstic acid $[\text{W}_4\text{O}_{13}\text{H}_2 + 8\text{H}_2\text{O} = \text{WO}_2(\text{OH})_2 + 3\text{WO}_3 + 8\text{H}_2\text{O}]$: Used for precipitating alkaloids.

Sodium tungstate $(\text{Na}_2\text{WO}_4 + 2\text{H}_2\text{O})$: Used for rendering fabric incombustible. Reagent for albumen.

W forms compound acids with P and with S .

$\text{WO}(\text{OH})_4$ was discovered by Scheele (1781). W was isolated some years later by the brothers De Luyart.

LXXIII. URANIUM.

Symbol, Ur . Hexad. Atomic weight, 238.482. Specific gravity, 18.38.

Occurs as oxide in *pitchblende*, as *uranium ochre*, *uranium mica*, etc. Not reducible by H but by heating UOCl_4 with Na . Metal resembles Fe , turns yellow in air, burns to black UO_2 . Soluble in dilute acids.

Does not decompose H_2O . Discovered by Klaproth (1789). Isolated by Peligot (1841).

Uranium tetrachloride (UCl_4): Greenish black regular deliquescent octohedra. Volatile at red heat.

Uranium pentachloride (UCl_5): Brownish-yellow. Volatile and deliquescent.

Uranium oxides: UO_2 , black, soluble in HNO_3 and also in concentrated H_2SO_4 ; UO_3 , yellow; U_3O_8 , green, native in *pitchblende*; U_2O_5 , black, formed by heating U_3O_8 .

Uranium hydroxide [$\text{U}(\text{OH})_4$]: Formed as brown precipitate by KOH .

Uranic acid [$\text{UO}_2(\text{OH})_2$]: Salts are also derived from $\text{U}_2\text{O}_7(\text{OH})_2$ as potassium uranate, $\text{K}_2\text{U}_2\text{O}_7$, and bismuth uranate ($\text{BiO})_2\text{U}_2\text{O}_7 + 3\text{H}_2\text{O}$ (native as uranosphærite).

Peruranic acid (UO_4): Also with $4\text{H}_2\text{O}$. Sodium peruranate forms as $\text{UO}_5\text{Na}_4 + 8\text{H}_2\text{O}$.

Uranous sulphate [$\text{U}(\text{SO}_4)_2 + 8\text{H}_2\text{O}$]: Green prisms in *johannite* (vitriol with Cu).

Uranous phosphate is green and gelatinous.

Uranic nitrate [$(\text{UO}_2)(\text{NO}_3)_2 + 6\text{H}_2\text{O}$]: Soluble in H_2O or alcohol.

Uranic carbonate [$(\text{UO}_2)\text{CO}_3$]: Forms double salts.

Uranic sulphate [$(\text{UO}_2)\text{SO}_4 + 3\text{H}_2\text{O}$]: Yellow.

Uranic phosphates: $(\text{UO}_2)_3(\text{PO}_4)_2$ not yet made; $\text{UO}_2\text{HPO}_4 + 4\text{H}_2\text{O}$, yellow, insoluble in H_2O ; $\text{UO}_2\text{H}_4(\text{PO}_4)_2 + 3\text{H}_2\text{O}$, yellow, soluble

Uranium sulphide (U_2S_3): Dark-gray and amorphous. Made in dry way.

Uranium oxysulphide (UO_2S): Chocolate brown; precipitated by $(\text{NH}_4)_2\text{S}$. Decomposes by heating to 40°C .

Uranous salts are green and *uranic* yellow.

Uranium glass fluoresces green.

LXXIV. TIN GROUP.

	Tl.	Zr.	Sn.	Th.
Atomic weight	48.0	89.367	117.498	238.414
Specific gravity	5.3	4.150	7.290	7.700

All tetrads.

Allied to silicium.

With halogens they form: mCl_2 ; m_2Cl_6 , and mCl_4 .

LXXV. TIN.

Symbol, Sn. (Stannum). Tetrad. Atomic weight, 117.698. Specific gravity, 7.7.

Occurs mostly as *tinstone* (SnO_2) in Cornwall, Malacca, Borneo,

Mexico, Saxony, and Bohemia. Also as *tin* (Sn). A white, lustrous metal, softer than Au and harder than Pb. Crystals are quadratic. Very ductile and malleable (tin-foil). Becomes brittle at 200° C. Long-keeping at low temperature converts it into a gray *brittle* modification of only 5.8 specific gravity, which may be broken with the fingers. When this is heated alone or with H₂O the specific gravity becomes 7.29 and the white color returns.

Melts at 228.5° C. Oxidizes in air. Boils at 1,450° to 1,600° C. Burns with white light. At ordinary temperatures does not oxidize. Dissolves with evolution of H in hot HCl; in H₂SO₄ with evolution of SO₂, forming SnSO₄; concentrated HNO₃ oxidizes it without solution; dilute HNO₃ slowly dissolves it to Sn(NO₃)₂ and with formation of NH₃; touching with Pt renders it passive against HNO₃.

Stannous chloride (SnCl₂): Is monoclinic with 2H₂O. Soluble. May be distilled at 617° to 628° C. A great reductant (Hg₂Cl₂ to HgCl₂ and Hg). In solution attracts oxygen, forming stannic chloride and oxychloride.

Stannic chloride (SnCl₄): Colorless liquid not solidifying at -20° C. Boils at 114° C. Fumes in air and forms *tin butter* (SnCl₄ + 3H₂O).

Double salts form; also *bromides*, *iodides*, and *fluorides*.

Stannous oxide (SnO): Brownish-black powder.

Sn₂O₃ is also a brownish-black powder.

Stannic oxide (SnO₂): Native as *quadratic tinstone*. Artificial is a white infusible powder, almost insoluble in acids. Fused with caustic alkalies it forms *stannates*.

Stannous hydrate [Sn(OH)₂]: Made by precipitating SnCl₂ with Na₂CO₃. White. Insoluble in H₂O or NH₄OH, but soluble in KOH, decomposed by boiling.

Stannic acid (H₂SnO₃?): Formed as white precipitate in SnCl₄ by Na₂Cl₃. Composition not definitely known. Soluble in HCl and in HNO₃; also in KOH and in NaOH.

Metastannic acid [Sn(OH)₄]: Formed by metallic tin and HNO₃. Other formula is: H₁₀Sn₅O₁₅ = 2SnO₂ + 2H₄SnO₄ + H₂SnO₃.

Stannous sulphate (SnSO₄): White, soluble in 5½ parts H₂O. *Double salts* form with K.

Stannous phosphate (SnHPO₄): White and granular.

Stannic sulphate [Sn(SO₄)₂]: White, easily decomposed into hydrate and acid.

Stannic phosphate [SnH(PO₃)₂]: White. Insoluble.

Potassium stannate (K₂SnO₃ + 3H₂O): Made by dissolving the hydrate in KOH. Colorless. Hexagonal or monoclinic.

Sodium stannate (Na₂SnO₃): Used as a mordant in calico printing.

Stannous sulphide (SnS): Brownish-black. Soluble in polysulphides.

Stannic sulphide (SnS_2): Crystallized in dry way or a yellow precipitate from stannic salts.

Tin alloys: Bronze; Sn and Cu. Pewter; Sn and Pb. Britannia; Sn and Sb. Mirror metal; Sn and Hg.

Solder melting at 194°C Sn 5 parts and Pb 1 part.
Solder melting at 260°C Sn 1 part and Pb 2 parts.
Solder melting at 182°C Sn 2 parts and Pb 1 part.

Tests for stannous salts:

1. Zn precipitates metallic Sn from solutions of SnCl_2 .
2. NaOH produces white precipitate of Sn(OH)_2 soluble in excess.
3. NH_4OH produces white precipitate of Sn(OH)_2 insoluble in excess.
4. H_2S or $(\text{NH}_4)_2\text{S}$ precipitates from acid or neutral solution dark-brown SnS .

Tests for stannic salts:

1. NaOH produces white precipitate from aqueous solution, soluble in excess. The NH_4OH precipitate is insoluble in excess.
 2. HgCl_2 produces no precipitate.
 3. H_2S produces yellow precipitate soluble in $(\text{NH}_4)_2\text{S}$.
- The *purple of cassius* is precipitated from a mixture of SnCl_2 and SnCl_4 by chloride of gold.

LXXVI. TITANIUM.

Symbol, Ti. Atomic weight, 48. Specific gravity, 5.3.

A rare metal discovered by Gregor (1791). Occurs as native oxides; *rutile*, *anatase*, *brookite*; also as titanate of iron. The metal is gray and burns with a brilliant light. Soluble in HCl or H_2SO_4 with evolution of H.

Titanous chloride (TiCl_2): Light, black powder. Decomposes H_2O and inflames.

Titanic tetrachloride (TiCl_4): Colorless liquid similar to SnCl_4 .

Titanic hexachloride (Ti_2Cl_6): Violet deliquescent scales.

Titanium also forms *bromides*, *iodides*, and *fluorides*.

Titanous oxide (Ti_2O_3): Black powder.

Titanic oxide (TiO_2): Occurs native as quadratic *rutile* and *anatase*; rhombic *brookite*. Artificial is a white infusible powder, nearly insoluble in acids.

Titanic acid [Ti(OH)_4]: Also TiO(OH)_2 . Both are similar to silicic acid.

Titanous sulphide [$\text{Ti}_2(\text{SO}_4)_3 + \text{H}_2\text{O}$] and *titanic sulphide* [$\text{Ti(SO}_4)_2 + 3\text{H}_2\text{O}$].

Titanium sulphide (TiS_2): Yellow scales.

Titanium and nitrogen unite as: TiN ; Ti_3N_4 ; Ti_5CN_4 .

LXXVII. ZIRCONIUM.

Symbol, Zr. Tetrad. Atomic weight, 89.367. Specific gravity, 4.15.

Occurs in *zircon*, *hyacinth*, and rare metals. Metal is black, amorphous, lustrous scales similar to Sb. Fusible with difficulty. Burns only before the compound blow-pipe. Oxidizes by burning with KOH. Soluble in HF.

Zirconium chloride (ZrCl_4): White, volatile, soluble in H_2O .

Zirconium oxychlorides: $\text{ZrOCl}_2 + 8\text{H}_2\text{O}$ and ZrOCl_6 .

Zirconium bromide (ZrBr_4) and *zirconium oxybromide* (ZrOBr_2).

Zirconium fluoride (ZrF_4): Also many double fluorides.

Zirconium oxide (ZrO_2): White, hard, infusible. Gives very intense light in oxyhydrogen flame.

Zirconium hydroxide: $\text{Zr}(\text{OH})_4$.

Zirconium sulphate [$\text{Zr}(\text{SO}_4)_2$]: Slightly soluble.

Zirconium silicate (ZrSiO_4): Native as quadratic *zircon* or *hyacinth*; with H_2O or as quadratic *malakone*. *Zirconates* form as: Na_2ZrO_3 .

Zirconium sulphide (ZrS_2): Only made in dry way. No precipitates in Zr solutions by H_2S . Alkaline sulphides precipitate the hydrate.

LXXVIII. THORIUM.

Symbol, Th. Tetrad. Atomic weight, 233.414. Specific gravity, 7.7.

Occurs as silicate in rare minerals; *thorite*, *orangite*. Also as phosphate with Ce, La, and Di, in *monacite*; also with Nb. Metal is dark-gray, burns in air, does not decompose H_2O . Soluble in HNO_3 , less in HCl, also in hot H_2SO_4 .

Thorium chloride (ThCl_4): Quadratic deliquescent plates.

Thorium fluoride: $\text{ThF}_4 + 4\text{H}_2\text{O}$. Also double fluorides, as $\text{ThF}_4 \cdot 2\text{KF} + 2\text{H}_2\text{O}$.

Thorium oxide (ThO_2): White infusible quadratic crystals.

Thorium hydroxide [$\text{Th}(\text{OH})_4$]: White. Soluble in acids; insoluble in alkalis.

Thorium sulphide (ThS_2): Similar to ZrS_2 .

Thorium nitrate, *sulphate*, and *silicate* also known.

LXXIX. YTTRIUM GROUP.

The members are related partially to the foregoing and partially to the *cerium group*. It contains a number of very rare metals, some as yet but imperfectly known. They are found in *gadolinite* and *orthite* as silicates; in *yttrotalite* and *euxenite* as tantalates and niobates, etc.

The oxide formerly known as *yttria* contains: Scandium [symbol, Sc; atomic weight, 43.98; ScO_3 , $\text{Sc}(\text{OH})_3$; $\text{Sc}(\text{SO}_4)_3 \cdot 3\text{K}_2\text{SO}_4$], thulium (Tu), yttrium (Yt, 89.816), erbium (Er, 165.891), holmium (Ho, 170), ytterbium (Yb, 172.76). Most of these form oxides (m_2O_3) as white or pink infusible powders. They have characteristic spectra.

LXXX. CERIUM GROUP.

	La.	Ce.	Di.	Sm.
Atomic weight.....	138.526	140.424	146.80	150.?
Specific gravity.....	6.100	6.620	6.54	

These rare metals are from the minerals *cerite*, *fluocerite*, *lanthanite*, *gadolinite*, *samaraskite*, etc. Sm has been separated from salts that were previously supposed to contain only Di. In 1885, by fractional crystallization Di salts were separated into green *praseodym* and red *neodym* salts. Both the flame and absorption spectra are peculiar.

LXXXI. CERIUM.

Symbol, Ce. Tetrad. Atomic weight, 140.424. Specific gravity, 6.62 to 6.72.

Discovered by Berzelius and Hiesinger (1803). Ductile, steel color, melts at red heat, oxidizes in moist air. Easily ignited by scraping with a knife, etc. Burns more brilliantly than Mg. Slowly decomposes H_2O . Soluble in dilute acids, but not in concentrated H_2SO_4 or HNO_3 .

Cerium chloride (CeCl_3): Made by burning Ce in Cl.

Cerium oxides (Ce_2O_3): White or grayish-blue powder and CeO_2 , salmon-colored powder or colorless regular octohedra. Soluble in H_2SO_4 .

Cerium hydroxide [$\text{Ce}(\text{OH})_3$ and $\text{Ce}_2\text{O}(\text{OH})_6$]: Yellow. Reagents for the alkaloids.

Cerium nitrate [$\text{Ce}(\text{NO}_3)_3 + 6\text{H}_2\text{O}$]: Pale red. Soluble. Used in medicine.

Cerium sulphate [$\text{Ce}(\text{SO}_4)_3 + 9\text{H}_2\text{O}$]: Rhombic octohedra. Soluble. Forms almost insoluble double salts with K or Na. When CeO_2 is dissolved in H_2SO_4 , brownish-red hexagonal $\text{Ce}_4(\text{SO}_4)_5 + 25\text{H}_2\text{O}$ and yellow $\text{Ce}(\text{SO}_4)_2 + 4\text{H}_2\text{O}$ are formed.

Cerium oxalate [$\text{Ce}_2(\text{C}_2\text{O}_4)_3 + 9\text{H}_2\text{O}$]: White. Used in medicine.

LXXXII. LANTHANUM.

Symbol, La. Tetrad. Atomic weight, 138.526. Specific gravity, 6.05 to 6.16.

Similar to Ce, but harder and less ductile. Oxidizes even in dry air. Burns brilliantly. Easily soluble in HNO_3 .

Lanthanum chloride (LaCl_3 and $\text{LaCl}_3 + 7\text{H}_2\text{O}$): Colorless prisms, very soluble in H_2O or alcohol.

Lanthanum oxide (La_2O_3): Formed by ignition of oxalate or nitrate. White powder, soluble in acids.

Lanthanum hydroxide [$\text{La}(\text{OH})_3$]: Formed as white gelatinous precipitate by KOH.

Lanthanic nitrate [$\text{La}(\text{NO}_3)_3 + 6\text{H}_2\text{O}$]: Colorless deliquescent tables.

Lanthanic sulphate [$\text{La}_2(\text{SO}_4)_3 + 9\text{H}_2\text{O}$]: Also anhydrous. Easily soluble. Forms little soluble double salts with K.

LXXXIII. DIDYMIUM.

Symbol, Di. Tetrad. Atomic weight, 146.80. Specific gravity, 6.54.

Yellowish and less fusible than Ce or La. Salts pink. Absorption spectrum has four dark lines between D and F, and two between F and G.

Latest investigations show didymium to consist of green *praseodym* and red *neodym*. The salts below of Di are as heretofore described.

Didymium chloride (DiCl_3): Pink. Soluble in H_2O . With $6\text{H}_2\text{O}$ is rose-red and monoclinic.

Didymium oxide (Di_2O_3): Has flame spectrum corresponding to absorption lines. White. On ignition forms DiO_2 (?).

Didymium nitrate [$\text{Di}(\text{NO}_3)_3 + 6\text{H}_2\text{O}$]: Deliquescent. Rose-red, with red solutions.

Didymium sulphate [$\text{Di}(\text{SO}_4)_3 + 8\text{H}_2\text{O}$]: Soluble. Rose-red. Monoclinic. Forms little soluble double salts with K, Na, NH_4 .

The separation of Di into its elements, by C. V. Wilsbach, is described as follows:

By repeated crystallization of a mixture of the double nitrates of La and Di with ammonium, the La salt was obtained pure, whilst the Di salt separated into the salts of two new elements, *neodymium* and *praseodymium*.

Neodymium, Nd, yields salts of a splendid amethyst color, which form rose-colored solutions. Its oxide when strongly ignited is blue; it dissolves readily in acids and behaves as a sesquioxide of its group.

Praseodymium, Pr, forms green salts. Its oxide, formed by the ignition of the nitrate, is dark-brown; it dissolves in H_2SO_4 with evolution of oxygen, and is readily reduced to a greenish-white sesquioxide when held in a reducing flame. The oxides of both elements, but especially that of praseodymium, possess the property of condensing oxygen on the surface. The atomic weights, 143.61 (for Pr), and 140.8 (for Nd), are provisionally assigned to the elements. Ta-

bles for the absorption spectra are given and described. When mixed in a certain proportion the elements give the original spectrum of Di.

LXXXIV. ANTIMONY GROUP.

	Vd.	Id.	Nb.	Sb.	Ta.	Bi.
Atomic weight	51.256	?	94.0	119.955	182.144	207.523
Specific gravity.....	5.5	?	7.6	6.710	7.020	9.820

Closely related to the nitrogen-phosphorus group. Members are pentatomic and triatomic, becoming more metallic with increase of atomic weight.

LXXXV. ANTIMONY.

Symbol, Sb (*stibium*). Triad and pentad. Atomic weight, 119.955. Specific gravity, 6.7. Melts, 430° C. Boils, 1,000° to 1,500° C.

Occurs rarely native as pure metal, mostly as sulphide. The metal is bluish white, not changed in air; when heated burns to white Sb_2O_3 . Crystals are octohedra, nearly cubical. Insoluble in HCl, soluble in hot concentrated H_2SO_4 , with liberation of SO_2 , oxidized by HNO_3 without solution. Dissolved by nitrohydrochloric acid.

Antimony alloys: With Pb, type metal; with Sn, britannia.

Antimony and hydrogen (SbH_3): Only known mixed with H. Made by action of H_2SO_4 on Zn and Sb. Colorless gas, similar to AsH_3 . Burns with greenish-white flame. Solidifies at -102.5°C . Gives mirror similar to AsH_3 , but of velvet-black color, and not soluble in solution of hypochlorite of sodium; also less volatile. Can not be made by Sb in alkaline solution with Zn (different from As as in Fleitmann's test).

Antimonous chloride (*butter of antimony*, SbCl_3): Colorless crystalline mass or rhombic crystals from CS_2 . Deliquescent.

Antimonous oxychloride (SbOCl): Made from SbCl_3 with 3 to 4 parts H_2O . White amorphous powder or monoclinic crystals. Soluble in CS_2 or chloroform; insoluble in alcohol or ether.

Algaroth's powder ($\text{Sb}_4\text{O}_5\text{Cl}_2$): Formed from SbOCl when heat drives off SbCl_3 . Also made from SbCl_3 by much H_2O .

Antimonic chloride (SbCl_5): Colorless liquid forming crystalline mass. Decomposed by H_2O . Antimony also forms *ous* and *ic bromides, iodides, and fluorides*.

Antimonous oxide (Sb_2O_3): Native as *senarmontite*, regular octohedra; or white *antimony glance*, rhombohedra. Formed with Sb_2O_4 by burning Sb in air or heating it with HNO_3 . Insoluble. Diamorphous like As_2O_3 . Vapor density at 1,500° C. = 19.9 = Sb_4O_6 .

Antimony hydroxide (SbO.OH): Dissolves in KOH or NaOH, but not in NH_4OH . Changed to Sb_2O_3 by boiling.

Antimonic oxide (Sb_2O_5): Pale yellow, amorphous. Insoluble.

Antimonic acid [$\text{SbO}(\text{OH})_3$]: White powder, slightly soluble in water, not soluble in HNO_3 .

Metantimonic acid [$\text{SbO}_2(\text{OH})$]: Made by heating $\text{SbO}(\text{OH})_3$.

Pyroantimonic acid [$\text{Sb}_2\text{O}_5(\text{OH})_4$]: Forms soluble K salts, but insoluble Na salts.

Antimonous sulphate [$\text{Sb}_2(\text{SO}_4)_3$]: Quadratic prisms. Decomposed by H_2O into $(\text{SbO})_2\text{SO}_4 + 2(\text{SbO}-\text{OH})$.

Antimonii et potassii tartaras, tartar emetic [$(\text{SbO})\text{KC}_4\text{H}_4\text{O}_6$] $_2 + \text{H}_2\text{O}$: Made by boiling Sb_2O_3 in cream of tartar. Rhombic octohedra. Soluble in 15 parts H_2O at 15°C ., or 8 parts at 100°C .

Antimonous sulphide (Sb_2S_3): Native as gray *antimony glance*, rhombic crystals, specific gravity, 4.7. Made by precipitating H_2S . Soluble in $(\text{NH}_4)_2\text{S}$, but not in $(\text{NH}_4)_2\text{CO}_3$.

Antimony oxysulphide ($\text{Sb}_2\text{S}_2\text{O}$): Native. Made artificially as *antimony cinnabar* by slowly heating at 80 to 90°C . tartar emetic with $\text{Na}_2\text{S}_2\text{O}_3$. *Antimony glance* is made by heating Sb_2S_3 . *Kermes' mineral*, official in 1870, is a mixture of about $2\text{Sb}_2\text{S}_3 + \text{Sb}_2\text{O}_3$. It is not a definite compound.

Antimonic sulphide (golden sulphuret, Sb_2S_5): Precipitated by H_2S from solution of $\text{SbO}(\text{OH})_3$ in HCl or from *Schlippe's salt* ($\text{Na}_3\text{SbS}_4 + 9\text{H}_2\text{O}$) by decomposition with acids. When heated it decomposes into S_2 and Sb_2S_3 .

Antimony sulphochloride (SbSCl_3): White, crystalline. Made from H_2S and SbCl_3 .

Tests for antimony salts:

1. H_2S produces orange precipitate in acid solutions, soluble in $(\text{NH}_4)_2\text{S}$.
2. Zn precipitates black Sb on platinum.
3. SbH_3 deposits velvety black spots on porcelain, which are insoluble in NaOCl .

Antimony was known to ancients as native sulphuret. Also in metalli state, as proven by a cast vase of pure metallic antimony found by Sarzec, in Tello (Persia), and ornaments from a necropolis in Transcaucasia, dating back to 4,000 years ago.

LXXXVI. BISMUTH.

Symbol, Bi. Triad and pentad. Atomic weight, 207.523. Specific gravity, 9.8 to 9.9. Melts, 264°C . Boils, 1,000 to 1,500 $^\circ \text{C}$.

Occurs mostly as native metal, rarely as sulphide (glance) or oxide, also with Te (*tetradymite*). Metal is white with pinkish luster, brittle at ordinary temperatures. Rhombohedra (almost cubical). Burns to yellow Bi_2O_3 . Insoluble in HCl , soluble in hot concentrated

H_2SO_4 with liberation of SO_2 , soluble in HNO_3 or nitrohydrochloric acid.

Bismuth tetrachloride (Bi_2Cl_4): Black, amorphous, hygroscopic, unstable.

Bismuth trichloride (BiCl_3): White deliquescent crystals. Forms double salts.

Bismuth oxychloride (BiOCl): White crystals. It is made by mixing BiCl_3 with H_2O . Splits into Bi_2O_3 and BiCl_3 by heating.

Bismuth also forms bromide (BiBr_3), iodide (BiI_3), and fluoride (BiF_3).

Bismuth dioxide (Bi_2O_2): Gray crystals. Not changed in dry air. Oxidized with H_2O .

Bismuth trioxide (Bi_2O_3): Yellow needles. Insoluble in H_2O or alkalis, but soluble in acids.

Bismuth hydroxide (BiOOH): White, amorphous. Insoluble in H_2O , but soluble in acids.

Bismuth oxygen salts: Either Bi replaces 3H of an acid, forming neutral salts, or BiO replaces 1H of acid, forming acid salts. Nearly all of the neutral salts are decomposed by H_2O into basic salts and acids.

Bismuth nitrate [$\text{Bi}(\text{NO}_3)_3 + 5\text{H}_2\text{O}$]: Made by dissolving Bi in HNO_3 . Large transparent crystals. Melts at 73°C ., and at 80°C . forms $2[\text{BiO}(\text{NO}_3)] + \text{H}_2\text{O}$. Soluble in glycerin. Also in glycerin with very little H_2O ; with much H_2O decomposes into $\text{Bi}(\text{NO}_3)(\text{OH})_2$.

Bismuth subnitrate [$\text{Bi}(\text{NO}_3)(\text{OH})_2$]: Formed by decomposing $\text{Bi}(\text{NO}_3)_3$ with much H_2O . When much washed there remains $\text{BiONO}_3 + \text{BiOOH}$.

Bismuth carbonate [$(\text{BiO})_2\text{CO}_3 \cdot \text{H}_2\text{O}$]: Official as subcarbonate. Made by precipitating $\text{Bi}(\text{NO}_3)_3$ with $(\text{NH}_4)_2\text{CO}_3$.

Bismuth sulphate [$\text{Bi}(\text{SO}_4)_3$]: Needles not easily decomposed by heat. When H_2SO_4 is added to $\text{Bi}(\text{NO}_3)_3$ it forms $\text{Bi} \cdot \text{BiO} \cdot (\text{SO}_4)_2 + 3\text{H}_2\text{O}$.

Bismuthic acid (HBiO_3): Made by passing Cl into solution of KOH with BiOOH in suspension. A chocolate-colored powder, turning red when washed with boiling H_2O . At 120°C . forms Bi_2O_5 .

Bismuth sulphide (Bi_2S_3): Native as *bismuth glance*. Artificially precipitated as a black powder by H_2S . Insoluble in alkaline sulphides, soluble in hot concentrated HCl and in HNO_3 .

Tests for bismuth salts:

1. If heated with KI and S on charcoal before the blow-pipe, a bright scarlet-red incrustation is formed.

2. Brownish-black Bi_2S_3 precipitated by H_2S or $(\text{NH}_4)_2\text{S}$. Insoluble in dilute acids or $(\text{NH}_4)_2\text{S}$.

3. Alkaline hydrates precipitate BiOOH . Insoluble in excess.

LXXXVII. VANADIUM.

Symbol, Vd. Pentad and triad. Atomic weight, 51.256. Specific gravity, 5.5.

Occurs as *lead vanadate* and in some ores as *valborhite*, *mothramite*, *purcherite*, etc. Metal is difficult to free from O. Is gray, powder not changed by air or H_2O . Burns to Vd_2O_5 . Insoluble in HCl, soluble in HF and in hot concentrated H_2SO_4 ; forms blue solutions with HNO_3 .

Vanadium dichloride ($VdCl_2$ or Vd_2Cl_4): Green and hygroscopic, forming violet solutions.

Vanadium trichloride ($VdCl_3$): Peach-bloom colored deliquescent tables, forming brown solutions which turn green on addition of HCl.

Vanadium tetrachloride ($VdCl_4$): Brown liquid, decomposed by H_2O and turns blue.

Vanadium oxychloride ($VdOCl_3$): Yellow liquid. Other oxychlorides also form.

Vanadium monoxide (VdO): Gray, lustrous insoluble powder. Soluble in acids. Absorbs O. Hydrate is brown.

Vanadium trioxide (Vd_2O_3): Black powder, insoluble in acids.

Vanadium dioxide (VdO_2): Blue crystalline powder. Hydrate is dark green.

Vanadium oxysulphate ($VdOSO_4$): Bluish-green powder.

Potassium hypovanadate ($Vd_4O_9K_2 + 7H_2O$): Reddish-brown scales.

Vanadic anhydride (Vd_2O_5): Reddish-brown rhombic prisms, soluble in acids. With reducents, turns first blue then green.

Vanadic acids: Pyro ($H_4Vd_2O_7$); meta ($HVdO_3$); ortho (H_3VdO_4). The ammonium salts are used for indelible inks.

Vanadium sulphide (VdS): Black. Insoluble in hot HCl or cold H_2SO_4 , soluble in hot concentrated H_2SO_4 , or in HNO_3 ; also in alkaline sulphides. Vd_2S_3 and Vd_2S_5 also form.

Vanadium nitride (VdN): Grayish-brown powder.

Tests for vanadium salts:

1. Are not precipitated from acid solutions by H_2S , but the solution in alkaline sulphides is precipitated by acids as brown oxysulphide soluble in alkaline sulphides or hydrates.

Vd was discovered by Sefström (1830).

LXXXVIII. NIOBIUM.

Symbol, Nb. Triad and pentad. Atomic weight, 94. Specific gravity, 7.06.

Rare minerals containing Ta usually contain Nb. Discovered by

Hatchet and named columbium. The metal is steel gray, burns in air. Insoluble in HCl or HNO_3 or aqua regia; soluble in hot concentrated H_2SO_4 .

Niobium chlorides: NbCl_3 ; NbCl_5 ; NbOCl_3 .

Niobium fluorides: NbF_5 ; NbF_5O ; and many double fluorides.

Niobium monoxide (Nb_2O_2): Black powder.

Niobium dioxide (Nb_2O_4): Black.

Niobium pentoxide (Nb_2O_5): White when cold, yellow while hot.

Niobium double salts: $\text{K}_8\text{Nb}_6\text{O}_{19} + 16\text{H}_2\text{O}$; $\text{KNbO}_3 + 2\text{H}_2\text{O}$; $2\text{K}_4\text{Nb}_2\text{O}_7 + \text{Nb}_2\text{O}_5 + 16\text{H}_2\text{O}$.

LXXXIX. TANTALUM.

Symbol, Ta. Triad and pentad. Atomic weight, 182.144. Specific gravity, 7.02.

This metal is a black powder, little known.

Tantalum pentachloride (TaCl_5): Yellow deliquescent crystals.

Tantalum pentabromide (TaBr_5), *tantalum pentafluoride* (TaF_5), and *double fluoride* ($\text{TaF}_5 + 2\text{KF}$) are known.

Tantalum dioxide (Ta_2O_4): Brown and insoluble.

Tantalum pentoxide (Ta_2O_5) and *tantallic acid* ($\text{H}_5\text{Ta}_6\text{O}_{19}$) are similar to the corresponding Nb compounds.

XC. IDUNIUM.

Symbol, Id. Discovered by Websky, in South American Vd ores (1884).

XCI. SILVER GROUP.

	Cu.	Ag.	Hg.	Pb.
Atomic weight.....	63.173	107.675	199.7120	206.471
Specific gravity.....	8.900	10.570	13.5953	11.370
Melts at.....	1,336° C.	1,000° C.	-39.5° C.	326° C.

XCII. LEAD.

Symbol, Pb (Plumbum). Tetrad (in organic compounds) and diad (in inorganic compounds). Specific gravity, 11.37.

Occurs rarely pure native, generally in ores; *galena* (PbS) or as carbonate, molybdate, phosphate, chromate, etc. *Made* from sulphide by roasting and smelting the product; or by smelting with Fe; $\text{Fe} + \text{PbS} = \text{FeS} + \text{Pb}$. In the roasting process $\text{PbS} + \text{O}_2 = \text{PbSO}_4$; add PbS and $\text{PbSO}_4 + \text{PbS} = 2\text{SO}_2 + 2\text{Pb}$. Also made by reducing PbO obtained by cupellation of argentiferous galena.

Metallic lead: Is bluish-gray, soft, and ductile. Melts at 326° C.,

volatile at high temperatures, oxidizes in air, burns to yellow oxide. With pure water it forms $\text{Pb}(\text{OH})_2$, somewhat soluble. Water containing Ca salts does not dissolve Pb. It is affected but little by HCl , HF , or H_2SO_4 , but HNO_3 dissolves it. Precipitated by Fe or Zn from solutions.

Lead chloride (PbCl_2): Soluble in 33 parts boiling or 135 parts cold H_2O . Crystals are colorless prisms. Vapor density is 9.5. Oxychlorides are formed by melting with PbO (Cassel yellow).

Lead tetrachloride (PbCl_4): Only known in solutions. Red, decomposed into Cl and PbO_2 by heat.

Lead bromide (PbBr_2): Rhombic crystals similar to PbCl_2 .

Lead iodide (PbI_2): Yellow scales. Almost insoluble in cold, but dissolves in 200 parts of boiling H_2O .

Lead fluoride (PbF_2): Also $\text{PbSiF}_6 + 2\text{H}_2\text{O}$ or $4\text{H}_2\text{O}$.

Lead suboxide (Pb_2O): Forms as a gray skin over melting Pb, made by heating oxalate of lead to 300°C . Black powder; at high temperature separates into Pb and PbO .

Lead monoxide (PbO): Litharge. Made by heating Pb or $\text{Pb}(\text{NO}_3)_2$ in air. Yellow or orange powder. Slowly absorbs CO_2 .

Lead sesquioxide (Pb_2O_3): Greenish-brown. Difficult to procure pure.

Red lead (Pb_3O_4): Made by carefully heating PbO in air to 400° or 500°C . Separated by acids into 2PbO and PbO_2 .

Lead dioxide (PbO_2): Heat separates it into O and PbO ; anhydride of a plumbic acid only known in salts; $\text{Pb}(\text{OH})_4$ or $\text{PbO}(\text{OH})_2$; $\text{K}_2\text{PbO}_3 + 3\text{H}_2\text{O}$.

Lead hydroxide [$\text{Pb}(\text{OH})_2$]: White. Seldom free from basic salt. Soluble in KOH or NaOH ; slightly soluble in pure H_2O , but insoluble in NH_4OH . Attracts CO_2 . $\text{Pb}_3\text{O}_2(\text{OH})_2$ is in octohedra.

Lead nitrate [$\text{Pb}(\text{NO}_3)_2$]: Octohedra. Soluble in 2 parts H_2O at 15°C . Insoluble in concentrated HNO_3 . Its solution in boiling H_2O dissolves PbO and forms $\text{PbONO}_3\text{H} = \text{NO}_2 - \text{O} - \text{Pb} - \text{OH}$.

Lead nitrite [$\text{Pb}(\text{NO}_2)_2\text{H}_2\text{O}$]: Formed by boiling Pb with $\text{Pb}(\text{NO}_3)_2$. Is yellow or red. Basic salts form containing the nitrite.

Lead carbonate (PbCO_3): Native as *cerusite* in rhombic silky crystals, or formed by precipitation of $\text{Pb}(\text{NO}_3)_2$ with $(\text{NH}_4)_2\text{CO}_3$. K_2CO_3 , etc., precipitates basic salts.

White lead [$2\text{PbCO}_3 + \text{Pb}(\text{OH})_2$]: Made on large scale by "corroding" thin "buckles" of Pb by alternate action of CO_2 and acetic acid vapor.

Phosgenite ($\text{PbCO}_3 + \text{PbCl}_2$): Yellow quadratic crystals.

Lead sulphate (PbSO_4): Native as *lead vitriol*. Insoluble in H_2O containing alcohol or in dilute H_2SO_4 . Soluble in concentrated H_2SO_4 , in solutions of sodium hyposulphite, ammonium tartrate, and other ammonium salts.

Lead chromate (PbCrO_4): Native as red lead ore. The artificial (chrome yellow) is precipitated from lead solutions by K_2CrO_4 . Insoluble in H_2O or dilute HNO_3 . Soluble in KOH or concentrated HNO_3 .

Lead dichromate (PbCr_2O_7): Red. Made by digesting PbCrO_4 with chromic acid. Forms basic red $\text{PbCrO}_4 + \text{PbO}$, when digested with KOH , or by adding PbCrO_4 to melted KNO_3 .

Lead phosphate [$\text{Pb}_3(\text{PO}_4)_2$]: The pyro and meta salts also form.

Lead arsenate [$\text{Pb}_3(\text{AsO}_4)_2$]: Other salts form similar to the phosphates. The native *mimesite* is $\text{Pb}_3(\text{AsO}_4)_2 + \text{PbCl}_2$.

Lead arsenite [$\text{Pb}(\text{AsO}_2)_2$]: Also PbHAsO_3 and $\text{Pb}_3(\text{AsO}_3)_2$.

Lead silicate: This is known in flint glass, enamels, etc.

Lead sulphide (PbS): Native as *galena* in cubes or octohedra. Made by precipitating acid or alkaline lead solutions with H_2S . There are many compounds of PbS with other sulphides: CuS , 2PbS .

Lead alloys: Type metal— Pb , 4 parts; Sb , 1 part. Pewter— Pb , 20 parts; Sn , 80 parts. Pb occurs in poor tin foil, tin plates, etc.

Tests for lead salts:

1. May be reduced on charcoal before the blow-pipe.
2. Precipitated by H_2S or $(\text{NH}_4)_2\text{S}$ from either neutral or acid solutions as black PbS , insoluble in acids.
3. $\text{K}_2\text{Cr}_2\text{O}_7$ precipitates yellow PbCrO_4 , insoluble in dilute HNO_3 , but soluble in NaOH .
4. Tincture of cochineal produces even in very dilute solutions a grayish-blue color.
5. H_2SO_4 forms insoluble precipitate of PbSO_4 .

XCIII. COPPER.

Symbol, Cu (cuprum). Diad. Specific gravity, 8.9. Occurs native in pure state as cubes; oxide as *copper ore*; sulphide as *copper glance*, with Fe and S in copper pyrites, etc. There are various methods of extracting the metal from ore.

Metallic Cu is red, very malleable and ductile, melts at $1,330^\circ \text{C}$., oxidizes slightly in air, with moisture and CO_2 forms basic carbonate, when heated in air forms black oxide; insoluble in dilute HCl or dilute H_2SO_4 if air be excluded, but in air and in contact with Pt or other metals Cu is slowly soluble; soluble in hot concentrated H_2SO_4 with evolution of SO_2 , readily soluble in HNO_3 , precipitated from solutions by Zn , Fe , P , etc.

Copper hydride (Cu_2H_2): Formed by adding concentrated solution of CuSO_4 to hypophosphorous acid and heating not above 70°C .; yellowish-brown and unstable. Decomposed at 60°C . into Cu and H .

Cuprous chloride (Cu_2Cl_2): Metallic Cu ignites in Cl , forming Cu_2Cl_2

in tetrahedra, melting at 434°C .; vapor density at $1,000^{\circ}\text{C}$. is 6.9. Colorless, turning violet and brown in sunlight. Insoluble in H_2O , but dissolves in $\text{Na}_2\text{S}_2\text{O}_3$, NH_4OH , or HCl . Solutions absorb hydrocarbons, etc. Compounds form, as $\text{Cu}_2\text{Cl}_2 + \text{CO} + 2\text{H}_2\text{O}$ (colorless); $\text{Cu}_2\text{Cl}_2 + 4\text{KCl}$ (octohedra); etc.

Cupric chloride (CuCl_2): Yellowish brown and deliquescent. Soluble in H_2O or alcohol. Solutions are green when concentrated, and blue when dilute. $\text{CuCl}_2 + 2\text{H}_2\text{O}$ form as pale blue rhombic prisms.

Oxychloride occurs native as *atacamite*.

Cuprous bromide (Cu_2Br_2) and *cupric bromide* (CuBr_2).

Cuprous iodide (Cu_2I_2): White.

Cuprous fluoride (Cu_2F_2) and *cupric fluoride* (CuF_2).

CuSiF_6 and CuCN also form.

Cuprous oxide (Cu_2O): Native in octohedra as red copper ore. Reduced from CuO by glucose.

Cuprous hydrate [$\text{Cu}_2(\text{OH})_2$]: Yellow as in Trommer's test.

Cupric oxide (CuO): Native as *tenorite*. An oxidizer in organic analysis.

Cupric hydrate [$\text{Cu}(\text{OH})_2$]: Blue. Loses H_2O by boiling.

There are but few staple cuprous salts with the oxygen acids.

Cuprous sulphite (Cu_2SO_3) and cuprous hyposulphite ($\text{Cu}_2\text{S}_2\text{O}_3$): Form double salts with the alkaline sulphites.

Cupric nitrate [$\text{Cu}(\text{NO}_3)_2 + 3\text{H}_2\text{O}$]: Dark-blue prisms. At low temperatures has $6\text{H}_2\text{O}$ and is in pale blue tables. Soluble in alcohol. Deliquescent and decomposed on heating to basic salts and finally to CuO . Double salts form as $\text{Cu}(\text{NO}_3)_2 + 4\text{NH}_3$; rhombic.

Cupric carbonate [$(\text{CuOH})_2\text{CO}_3$]: Native as *green malachite*. The neutral carbonate is unknown. Copper lazur-blue is $\text{Cu}(\text{CuOH})_2(\text{CO}_3)_2$. Forms double carbonates with alkaline carbonates. Forms as *verdigris* on Cu and brass vessels (true verdigris is a basic acetate).

Cupric sulphate ($\text{CuSO}_4 + 5\text{H}_2\text{O}$, blue vitrol): Manufactured from native sulphide by roasting, etc. Triclinic. Soluble in $2\frac{1}{2}$ parts of H_2O at 15°C . or $\frac{1}{2}$ part at 100°C . The anhydrous salt is white, used as reagent for H_2O . Various basic salts. With alkaline sulphates forms monoclinic compounds, isomorphous with those of Mg or Zn having $6\text{H}_2\text{O}$.

Cuprammonium sulphate [$\text{CuSO}_4 + 2\text{NH}_3 = (\text{N}_2\text{H}_6\text{Cu})\text{SO}_4$]: Also forms with $2\text{NH}_4\text{OH}$ additional.

Cupric phosphate [$\text{Cu}_3(\text{PO}_4)_2 + 3\text{H}_2\text{O}$]: Bluish-green. Soluble in acids or NH_4OH .

Cupric arsenate (CuHAsO_4): Greenish-blue precipitate. Also native.

Cupric arsenite (CuHAsO_3): Scheele's green. Made by precipitation of CuSO_4 with solution of arsenous acid while carefully neutralized. Insoluble in H_2O , soluble in KOH . Blue.

Cupric silicate ($\text{CuSiO}_3 + \text{H}_2\text{O}$): *Diopase*. Hexagonal. Green.

Cuprous sulphide (Cu_2S): Native as *copper glance*. Made by burning Cu in S vapor. Also by action of NH_4HS on Cu. Dark-gray, rhombic, and regular. Double salts as native *copper-silver glance* ($\text{Cu}_2\text{S} + \text{Ag}_2\text{S}$); CuSbS_2 ; Cu_3BiS_3 ; etc.

Cupric sulphide (CuS): Copper indigo. Hexagonal. Precipitated from acid solution of Cu salts by H_2S . Forms double alkaline sulphides.

Cu_2Se native as *berzelianite*.

Cu_3N_2 is a dark-green powder.

Cu_3P_2 is a black precipitate formed in Cu solutions by PH_3 .

Cu_3As is native in Chili.

Copper alloys:

Brass,	71.5 per cent Cu; 28.5 per cent Zn.
Tombac	84.5 per cent Cu; 15.5 per cent Zn.
White,	30.0 per cent Cu; 70.0 per cent Zn.
Bronze (ordinary),	70 per cent Cu; 30 per cent Sn.
Bronze (for guns),	90 per cent Cu; 10 per cent Sn.
Bronze (for bells),	80 per cent Cu; 20 per cent Sn.
Bronze (for mirrors),	66⅔ per cent Cu; 33⅓ per cent Sn.

Reitz's bronze to resist acids: Cu, 15; Zn, 2.34; Pb, 1.82; and Sb, 1 part.

Tests for copper:

1. Bright metallic iron precipitates Cu from solution of its salts.
2. NH_4OH precipitates $\text{Cu}(\text{OH})_2$, which re-dissolves in excess, forming blue solution.
3. KOH precipitates pale blue $\text{Cu}(\text{OH})_2$ from cupric solutions. Heat changes it to black CuO .
4. Potassium ferrocyanide forms brownish-red precipitate.
5. H_2S forms black insoluble CuS .
6. Cuprous salts are red or yellow; anhydrous cupric salts are white, and hydrous salts are blue or green.

Copper poisoning: Cu vessels used for cooking, for mineral waters, or acid liquids, or wines, often produce poisoning, shown by nausea and vomiting. The antidotes are milk, albumen, ferrocyanide of potassium, etc.

XCIV. MERCURY.

Symbol, Hg (Hydrargyrum). Diad. Atomic weight, 199.712. Specific gravity, 13.5953. Freezes, -39.5°C . Boils, 360°C . Synonym, quicksilver.

Occurs in metallic drops, or as *cinnabar* (HgS).

Obtained by distilling HgS with Fe or CaO. Rectified by redistillation, or by acids.

The metal is silver-white, regular octohedra below -39.5°C . Evap-

orates at ordinary temperatures. Vapor density is 6.9 (air=1). Unchanged at ordinary temperatures, but oxidizes near 360° C. Insoluble in HCl or in cold H_2SO_4 , but easily soluble in hot H_2SO_4 or in HNO_3 . The vapor is very poisonous. The molecule contains but one atom. Precipitated as metal from its solutions by other metals, as SnCl_2 , in gray powder. Exists in fine subdivision in blue mass, ointments, etc.

Forms mercuric and mercurous compounds. The latter contain the greater proportion of Hg.

Mercurous chloride (calomel, Hg_2Cl_2): Native as *horn-quicksilver*. Precipitated by HCl from mercurous salts. Made by subliming 4 parts HgCl_2 (corrosive sublimate) with 3 parts Hg and washing the product. Also made by adding 50 pounds Hg to 70 pounds H_2SO_4 , giving 62 pounds HgSO_4 ; to this $40\frac{1}{2}$ pounds Hg is added and well mixed with 34 pounds NaCl; 95 to 100 pounds Hg_2Cl_2 is obtained. Quadratic crystals. Specific gravity is 7.2. Decomposed by light. Vapor density at 440° C. is 8.21, due to dissociation of Hg_2Cl_2 ($\text{Hg} + \text{HgCl}_2$). Insoluble in H_2O , alcohol, or dilute acids. Incompatible with alkaline chlorides (?); soluble iodides, bromides, acids; in short, all bodies tending to separate Hg_2Cl_2 into $\text{Hg} + \text{HgCl}_2$.

Mercuric chloride (corrosive sublimate, HgCl_2): Made by dissolving Hg in aqua regia; HgO in HCl; or by subliming HgSO_4 with NaCl. Rhombic prisms. Specific gravity is 5.4. Melts at 288° C. Soluble in 14 parts H_2O at 15° C., or 2 parts at 100° C. Easily soluble in alcohol. Very poisonous. Antiseptic. Forms double chlorides and oxychlorides; *mercurammonium* ($\text{HgCl}_2 + 2\text{NH}_4\text{Cl} + \text{H}_2\text{O}$).

Mercurous bromide (Hg_2Br_2) and *mercuric bromide* (HgBr_2) are similar to the corresponding chlorides.

Mercurous iodide (green iodide, Hg_2I_2): Greenish yellow. Made by intimately triturating Hg_2 with I_2 , or by Hg with HgI_2 , or by cautious addition of KI to mercurous salts. Very little soluble in H_2O , insoluble in alcohol. Decomposed by light, by KI, etc.

Mercuric iodide (red iodide, HgI_2): Made by direct union of Hg and I_2 or by HgCl_2 and $(\text{KI})_2$. Insoluble in H_2O . Yellow or red. Sublimes in yellow, rhombic crystals, turning red on contact. Double salts form as: HgICl (yellow), HgIBr (yellow).

Mercury and iron: Hg_2Fe_2 and HgFe_2 .

Mercuric cyanide [$\text{Hg}(\text{CN})_2$]: White, quadratic. On heating, yields CN and Hg. Double cyanides form.

Mercurous oxide (Hg_2O): Formed from Hg_2Cl_2 and KOH (black wash). Decomposed by light and heat.

Mercuric oxide (HgO): Amorphous is yellow and the crystalline red. At red heat is decomposed into Hg and O. Easily soluble in acids.

Mercurous nitrate [$\text{Hg}_2(\text{NO}_3)_2 + 2\text{H}_2\text{O}$]: Made by dissolving excess of Hg in cold HNO_3 . Monoclinic yellow basic salt is separated by H_2O ; $\text{Hg}_2(\text{NO}_3)_2 + \text{Hg}_2\text{O} + \text{H}_2\text{O}$.

Mercurous carbonate (Hg_2CO_3): Precipitated from solutions of mercurous salts by K_2CO_3 . Yellow and unstable.

Mercurous sulphate (Hg_2SO_4): White. Monoclinic. Little soluble in cold and decomposed by boiling H_2O .

Mercurous phosphate [$\text{Hg}_6(\text{PO}_4)_2$]: White amorphous precipitate by Na_2HPO_4 .

Mercuric nitrate [$\text{Hg}(\text{NO}_3)_2$]: Formed by dissolving Hg in hot HNO_3 . Rhombic deliquescent crystals. Used in determining urea. The basic salts are: $\text{Hg}(\text{NO}_3)_2 \cdot \text{HgO} + 2\text{H}_2\text{O}$ and $\text{Hg}(\text{NO}_3)_2 \cdot 2\text{HgO} + \text{H}_2\text{O}$. They lose HNO_3 by boiling and leave HgO .

Mercuric carbonate ($\text{HgCO}_3 + 3\text{HgO}$): Precipitated from mercuric solutions by Na_2CO_3 .

Mercuric sulphate (HgSO_4): White; yellow on heating. With much H_2O forms basic *Turpethum minerale* ($\text{HgSO}_4 + 2\text{HgO}$). Double salts form as $3\text{HgSO}_4 + \text{K}_2\text{SO}_4 + 2\text{H}_2\text{O}$.

Mercuric phosphate [$\text{Hg}_3(\text{PO}_4)_2$]: Precipitated from solutions of mercuric nitrate (not from chloride) by Na_2HPO_4 . White. Insoluble in H_2O .

Mercurous sulphide (Hg_2S): Black. Very unstable ($\text{HgS} + \text{Hg}$).

Mercuric sulphide (HgS): Native as hexagonal *cinnabar*. Precipitated amorphous by H_2S . Black, but red after sublimation or by trituration with an alkaline sulphide. Insoluble in HCl or in HNO_3 , but soluble in aqua regia, in K_2S , or Na_2S (if KOH be present), but not in $(\text{NH}_4)_2\text{S}$. Double salts form as $\text{HgK}_2\text{S}_2 + 5\text{H}_2\text{O}$; $\text{HgNa}_2\text{S}_2 + 5\text{H}_2\text{O}$. If H_2S be added to excess of HgCl_2 , $2\text{HgS} + \text{HgCl}_2$ forms. Other compounds form with Br, I, or F.

Mercuric nitride (Hg_3N_2): Dark brown. Very explosive.

Mercurous-ammonium chloride ($\text{NH}_2\text{Hg}_2\text{Cl}$): Black. Formed by treating Hg_2Cl_2 with NH_4OH .

Mercurous-ammonium nitrate ($\text{NH}_2\text{Hg}_2\text{NO}_3$) also forms.

Mercurammonium chloride (NH_2HgCl): Infusible white precipitate. When boiled it forms $\text{O} \begin{matrix} \text{HgNH}_2 \\ \text{HgCl} \end{matrix}$.

Mercuri-diammonium chloride ($\text{N}_2\text{H}_6\text{HgCl}_2$): Formed by adding HgCl_2 to a boiling mixture of $\text{NH}_4\text{Cl} + \text{NH}_4\text{OH}$. Regular rhombicdodecahedra. The official *white precipitate* is a mixture of this and NH_2HgCl .

Mercurammonium hydroxide ($\text{NH}_2\text{HgOH} \cdot \text{HgO}$): Yellow powder. Explodes on heating.

Tetramercurammonium oxide ($\text{N}_2\text{Hg}_4\text{O}$): Explosive.

Mercury alloys (amalgams): Native regular crystals as AgHg . Hg forms amalgams with K, Na, Li, Ca, Ba, Sr, Zn, Cd, Bi, Au, Cu, etc. CuHg is used for filling teeth.

Tests for Mercury:

1. Forms metallic spots on Cu.
2. Solution of KI forms HgI_2 .

3. Sublimation with NaCO_3 forms metallic globules.
4. In acid solutions H_2S forms black HgS .

XCV. SILVER.

Symbol, Ag (argentum). Diad. Atomic weight, 107.675. Specific gravity, 10.57. Melts $1,000^\circ \text{C}$.

Occurs native as pure silver (wire silver), also as chloride (hornblende), bromide, and iodide. Occurs in small quantity in galena.

Obtained from galena by passing it into the metallic Pb. The alloy at first lets the Pb crystallize, then the rich fused residue is cupellated. When obtained by the Zn process, Zn is added to the argentiferous Pb and forms an alloy with the Ag, which floats on the melted Pb. Is then purified by distilling off the Zn. Rich ores are sometimes amalgamated after roasting with NaCl.

Ag is alloyed for coins and vessels. Pure Ag obtained from the chloride. Occlusion of O in molten Ag.

Ag is a pure white metal in regular octohedra. Melts at $1,000^\circ \text{C}$. Unchanged in air unless S be present. Insoluble in HCl. Soluble in concentrated H_2SO_4 and in dilute HNO_3 ; also soluble in HI.

Silver chloride (AgCl): Native as *horn silver* in regular crystals. Precipitated by soluble chlorides from solutions of silver. White, but light turns it first violet, then black. Used in photography. Melts at 451°C . Insoluble in water or dilute acids, but soluble in concentrated alkaline chlorides, concentrated HCl, in NH_4OH , KCN, $\text{Na}_2\text{S}_2\text{O}_3$, and in concentrated solutions of $\text{Hg}(\text{NO}_2)_2$. Dry AgCl absorbs NH_3 , forming $\text{AgCl} + 3\text{NH}_3$.

Silver bromide (AgBr): Native in Mexico and Chili. Pale yellow. Less soluble in NH_4OH than AgCl .

Silver iodide (AgI): Yellow, amorphous. Ag dissolves in HI with evolution of H and forming AgI.HI , which with excess of Ag forms AgI . Insoluble in H_2O or NH_4OH but soluble in KCN, in concentrated solution of KI, in $\text{Na}_2\text{S}_2\text{O}_3$. Decomposed by light only when impure (containing AgNO_3).

Silver fluoride ($\text{AgF} + 2\text{H}_2\text{O}$): Colorless prisms.

Silver cyanide (AgCN): Made from silver solutions by soluble cyanides or HCN. White. Curdy. Not decomposed by light but by heat. Soluble in KCN, etc. $\text{AgCN} + \text{KCN}$ is in hexagonal plates. Used in silver plating.

Silver monoxide (Ag_2O ; $\begin{array}{c} \text{Ag} \\ | \\ \text{Ag} \end{array} \text{O}$): Dark brown. Amorphous. Slightly soluble in H_2O . Decomposed at 250°C . Reduced at 100°C . by H. Ignites finely powdered S, Sb_2S_3 , P, etc.

Silver peroxide (Ag_2O_2 ; $\begin{array}{c} \text{Ag}-\text{O} \\ | \quad | \\ \text{Ag}-\text{O} \end{array}$): Made from action of ozone on Ag or electricity on AgNO_3 . Black. Easily decomposed. Ignites S, H_2S , etc.

Silver suboxide (Ag_4O ; $\begin{array}{c} \text{Ag}-\text{Ag} \\ | \quad | \\ \text{Ag}-\text{Ag} \end{array} \text{O}$): Made by heating $\text{Ag}_2\text{C}_6\text{H}_5\text{O}_7$ in hydrogen. Very unstable. Black.

The only salts of Ag with O acids known are those in which Ag replaces H, as the solutions of AgO or Ag_2O_2 in acids decompose by heat.

Silver nitrate (AgNO_3): Colorless, rhombic tables, of 4.3 specific gravity. Melts at 218°C . Soluble in $\frac{1}{2}$ part H_2O at 15°C .; less soluble in HNO_3 , and slightly soluble in alcohol. Neutral reaction. When pure, is unchanged in light or air. Reduced by organic substances. Blackens. Poisonous. Strong caustic (lunar caustic). Various dilutions in commerce. Its concentrated hot solutions dissolve haloid Ag salts, forming compounds decomposed by H_2O . In powder, AgNO_3 absorbs 3NH_3 , and from aqueous solutions of rhombic $\text{AgNO}_3 + 2\text{NH}_3$ forms.

Silver nitrite (AgNO_2): Precipitated from solutions of AgNO_3 by KNO_2 . Little soluble in cold H_2O , more so in hot H_2O . Decomposed at 90°C .

Silver chlorate (AgClO_3): Quadratic. Soluble in 5 parts H_2O . Detonates on heating.

Silver carbonate (Ag_2CO_3): Precipitated from Ag solutions by K_2CO_3 . Pale yellow. Insoluble in H_2O .

Silver sulphate (Ag_2SO_4): Rhombic. Little soluble in cold, but more so in hot H_2O . The acid sulphate (AgHSO_4) is in yellow prisms.

Silver sulphite (Ag_2SO_3): Formed by adding H_2SO_3 to AgNO_3 . White. Curdy. Decomposes.

Silver orthophosphate (Ag_3PO_4): Yellow. Amorphous. The acid salt (Ag_3HPO_4) is white and crystalline.

Silver pyrophosphate ($\text{Ag}_4\text{P}_2\text{O}_7$): White. Insoluble in H_2O . Easily soluble in HNO_3 or NH_4OH .

Silver metaphosphate (AgPO_3): White. Amorphous. Various modifications.

Silver arsenate (Ag_3AsO_4): Reddish brown. Amorphous. Insoluble in H_2O . Easily soluble in HNO_3 or NH_4OH .

Silver arsenite (Ag_3AsO_3): Yellow. Formed by adding As_2O_3 to AgNO_3 and a little NH_4OH . Decomposes on boiling into Ag, Ag_3AsO_4 , and As_2O_3 . By boiling with KOH it forms K_3AsO_4 , Ag, and Ag_4O .

Silver sulphide (Ag_2S): Native as regular *argentite* or rhombic *da-*

leminzite. Formed by H_2S even on metallic Ag in air, or from solutions.

Silver nitride (Ag_3N): Fulminating Ag. Formed by digesting Ag_2O in NH_4OH . Black crystals. Very dangerous explosive. $AgNH_2$ is questionable.

Silver phosphide (Ag_3P): Dark gray. Ag also forms Ag_3Sb , Ag_3Sb , and Ag_3Sb .

Silver alloys: Coin silver is Ag 90 parts and Cu 10 parts. Silverware is Ag 75 parts and Cu 25 parts.

Tests for Ag salts:

1. Soluble chlorides form white precipitate of $AgCl$.
2. H_2S forms black precipitate of AgS .
3. Reduced by Cu, Zn, etc.
4. Reduced from alkaline solutions by such organic substances as milk sugar, tartaric acid, formic acid, chloral hydrate, etc.

XCVI. GOLD.

Symbol, Au (aurum). Monad and triad. Atomic weight, 196.155 (Kruess in 1887 gives 196.640). Specific gravity, 19.33. Melts, $1,240^\circ C$.

Occurs mostly native as pure metal, generally associated with Ag, also with Te; in sand of many rivers as gold dust.

The metal is yellow, lustrous, and forms regular crystals. Softer than silver and more ductile and malleable. Melts to green liquid at $1,240^\circ C$. Not oxidized by O or H_2O . Insoluble in H_2SO_4 , HCl , or HNO_3 , but soluble in aqua regia. Reduced from solutions by Fe_2SO_4 , oxalic acid, etc. In thin laminæ it transmits green, blue, or violet light, according to thickness of the leaf.

Gold monochloride (aurous chloride, $AuCl$): Formed by heating $AuCl_3$ to $200^\circ C$. while stirring. White, insoluble powder.

Gold dichloride ($AuCl_2$): Formed by direct union of Cl and Au. Decomposes easily into $AuCl_3$.

Gold trichloride (auric chloride, $AuCl_3$): Formed by dissolving Au in aqua regia. Brown crystals. Deliquescent to dark red liquid. Sublimes in current of Cl. The hydrous ($AuCl_3 + 2H_2O$) as orange crystals. At $185^\circ C$. $AuCl_3$ splits into Cl_2 and $AuCl$. Many double salts, as $AuCl_3 + HCl + 4H_2O$ (auro-chloric acid), $AuCl_3 + KCl + 2\frac{1}{2}H_2O$ (yellow tables).

Gold bromides ($AuBr$, $AuBr_2$, and $AuBr_3$): Similar to the chlorides.

Other salts form, as: AuI , AuI_3 , $AuCN$, $AuCN + KCN$, $Au(CN)_2 + 3H_2O$, etc.

Aurous oxide (Au_2O): Formed by decomposing $AuCl$ with KOH . Decomposes at $160^\circ C$. Not soluble in HNO_3 or H_2SO_4 .

Auric oxide (Au_2O_3): Brown powder. Decomposes at 100°C .

Auric acid $[\text{Au}(\text{OH})_3]$: Brown.

Potassium aurate ($\text{KO}-\text{AuO}+3\text{H}_2\text{O}$): Yellow.

Gold and sodium hyposulphite ($\text{Au}_2\text{S}_2\text{O}_8 + 8\text{Na}_2\text{S}_2\text{O}_3 + 4\text{H}_2\text{O}$): Colorless.

Gold sulphide ($\text{Au}_2\text{S}_2?$): Formed from cold solutions of $\text{AuCl} + \text{H}_2\text{S}$ (from hot solution, Au falls). Black. Soluble in alkaline sulphides (forming double sulphides).

Fulminating gold: Formed by digesting the oxide (Au_2O_3) in ammonia. Formula not known.

Gold alloys: With Cu and Ag for coins and ornaments. The fineness is counted by *carats* (24 carats is pure gold). Fifty-eight per cent Au + 42 per cent Cu = 14 carats. Coins = 90 per cent Au + 10 per cent Cu (ducats contain 2 per cent Cu).

Tests for gold salts:

1. Solution of stannous with stannic chloride gives *purple of Cassius*.
2. Reduced to metallic state by ferrous sulphate, oxalic acid, arsenous acid, and other reducing agents.

XCVII. PLATINUM GROUP.

	Ru.	Rh.	Pd.	Os.	Ir.	Pt.
Atomic weight.....	104.217	104.055	105.737	198.494	192.651	196.155
Specific gravity.....	12.260	12.100	11.800	22.400	22.380	21.500

All occur together as native alloys (platinum ore) in Ural, California, Brazil, Colombia, Borneo, and Australia. Pt is the most abundant, as it forms from 50 to 80 per cent of the native alloy.

Uralium (?), U1 (atomic weight, 187.25; specific gravity, 20.25), and davydum, Da (atomic weight, 150.00; specific gravity, 9.39) also announced as elements, and form a heavy and a light group.

XCVIII. PLATINUM.

Symbol, Pt. Tetrad. Atomic weight, 196.155; specific gravity, 21.50. Melts, $1,800^\circ\text{C}$.

Discovered by Wood in 1741, and first described by Wollaston & Scheffer.

Has a grayish-white luster, very ductile, melts in oxyhydrogen flame, volatile. It may be welded like Fe at white heat. In powder as *platinum sponge*, and still finer as *platinum black*, it condenses H (absorbing more than 200 volumes), ignites H, and oxidizes alcohol. At red heat Pt is permeable by H, but not by O, N, or CO_2 . Oxidized by fusing KOH or KNO_3 . Insoluble in all acids except aqua regia.

Platinous chloride (PtCl_2): Formed by union of Pt and Cl_2 at $1,400^\circ \text{C.}$; only stable at higher or lower temperature (unstable at medium temperatures). Grayish green. Insoluble in H_2O . Soluble in hot HCl . Soluble quadratic crystals of double salts form with K or NH_4OH . Also made by heating PtCl_4 to 230°C.

Platinic chloride (PtCl_4): Formed by dissolving Pt in aqua regia. The crystals with $10\text{H}_2\text{O}$ are deliquescent and brown; with $5\text{H}_2\text{O}$ are red and not deliquescent. Soluble in H_2O or alcohol. With HCl forms $\text{PtCl}_4 \cdot 2\text{HCl} + 6\text{H}_2\text{O}$. Yellow, insoluble double salts form with K, Cs, Rb, or NH_4 . The orange triclinic Na salt is soluble. The Th salt is insoluble. Salts with Ca, Ba, Sr, Mg, or Zn are soluble, and after the type $\text{PtCl}_4 + \text{CaCl}_2 + \text{water}$. Soluble salts form with Al or Fe as $\text{PtCl}_4 \cdot \text{AlCl}_3 + \text{water}$.

Platinous bromide (PtBr_2) and *platinic bromide* (PtBr_4) are similar to the chlorides.

Platinous iodide (PtI_2) and *platinic iodide* (PtI_4) also similar to chlorides. $\text{PtI}_4 + 2\text{KI}$ is easily soluble.

Platinous cyanide [$\text{Pt}(\text{CN})_2$]: Greenish yellow. Insoluble in H_2O , alkalies, or acids. Many beautiful double cyanides form as $\text{Pt}(\text{CN})_2 + 2\text{HCN}$. $\text{Pt}_2(\text{CN})_6$ is only known in double salts.

Platinous oxide (PtO): Black powder.

Platinic oxide (PtO_2): Black powder. Insoluble in acids.

Platinosoplatinic oxide (Pt_3O_4): Black powder.

Platinous hydrate [$\text{Pt}(\text{OH})_2$]: Black powder. Soluble in dilute H_2SO_4 .

Platinic hydrate [$\text{Pt}(\text{OH})_4$]: Reddish-brown powder with acid properties. Soluble in NaOH or dilute acids.

Platinous nitrite [$\text{Pt}(\text{NO}_2)_2$]: Only known in double salts as; $\text{Pt}(\text{NO}_2)_2 \cdot 2\text{KNO}_2$.

Platinic sulphate: Has the formula $\text{Pt}(\text{SO}_4)_2$.

Platinous sulphite [$\text{Pt}(\text{SO}_3)$]: Forms double salts with K or Na.

Platinous sulphide (PtS): Formed by H_2S . Soluble in alkaline sulphide.

Platinic sulphide (PtS_2): Formed by H_2S . Soluble in alkaline sulphides.

Ammonium compounds of platinum: $\text{PtSO}_4(\text{NH}_3)_4$; $\text{Pt}(\text{NO}_3)_2 \cdot (\text{NH}_3)_4$; $\text{Pt}(\text{NH}_3)_4(\text{OH})_2$, etc.

Experiments with platinum: Döbereiner's lamp. Kindling of H with Pt sponge. Glowing Pt spiral in ether and air. Melting Pt in oxy-hydrogen flame.

XCIX. PALLADIUM.

Symbol, Pd. Tetrad. Atomic weight, 105.737; specific gravity, 11.8; melts, $1,800^\circ \text{C.}$

Discovered by Wollaston in 1803.

Whiter than Pt, very ductile and malleable, melts and volatilizes at slightly lower temperatures than Pt. Oxidizes superficially; reduced at white heat. In alcohol or gas flames Pd becomes sooty from absorption of H, of which it occludes 900 volumes (Pd_2H and other compounds richer in H). Soluble in concentrated HNO_3 and as fine powder in HCl or H_2SO_4 ; also in aqua regia.

Palladious chloride (PdCl_2): Dark brown and deliquescent. Reagent for I. Double salts form with K, etc.

Palladic chloride (PdCl_4): Only known in solutions.

Palladious iodide (PdI_2): Precipitated from solutions of iodide by PdCl_2 or $\text{Pd}(\text{NO}_3)_2$.

Palladious bromide (PdBr_2): Precipitated from solutions of bromides by $\text{Pd}(\text{NO}_3)_2$.

Palladium cyanide [$\text{Pd}(\text{CN})_2$]: Precipitated from solution of PdCl_2 by $\text{Hg}(\text{CN})_2$.

Palladium oxides: Pd_2O ; black powder. PdO ; black powder, formed by heating $\text{Pd}(\text{NO}_3)_2$. PdO_2 ; black powder.

Palladium hydrates: None isolated.

Palladious sulphate ($\text{PdSO}_4 + 2\text{H}_2\text{O}$): Brown soluble crystals.

Palladious nitrate [$\text{Pd}(\text{NO}_3)_2$]: Deliquescent.

PdS : Grayish white, or from H_2S : brownish black. PdS_2 and Pd_2S also form.

Palladium ammonium compounds: Similar to corresponding Pt compounds.

Tests for palladium salts:

1. Occludes hydrogen.
2. Brownish-black precipitate formed by H_2S .
3. Has black iodide.
4. Has yellow cyanide.

C. IRIDIUM.

Symbol, Ir. Tetrad. Atomic weight, 192.651. Specific gravity, 22.38. Melts, $1,950^\circ\text{C}$.

A metal similar to polished steel. Harder, less fusible, and more brittle than Pt. Insoluble in acids; except as a fine powder, in aqua regia. Oxidized by heating with KOH , KNO_3 , or KClO_3 . Discovered by Tennant in 1803.

Iridium dichloride (IrCl_2): Dark brown. Soluble in H_2O .

Iridium tetrachloride (IrCl_4): Dark brown. Soluble in H_2O . Forms double salts with K.

Iridium hexachloride (Ir_2Cl_6): Formed by heating Ir in Cl . Olive green. Insoluble.

Iridium monoxide (IrO): Black powder.

Iridium trioxide (Ir_2O_3): Black. Insoluble. Reduced by H in cold.

Iridium dioxide (IrO_2): Black. Insoluble.

Iridium hydroxides: Black $[\text{Ir}_2(\text{OH})_6]$ and blue $[\text{Ir}(\text{OH})_4]$.

CI. RHODIUM.

Symbol, Rh. Tetrad. Atomic weight, 104.055. Specific gravity, 12.1. Melts, $2,000^\circ \text{C}$.

A grayish-white metal. Melts in oxyhydrogen flame. Ductile. Not volatile. Dissolves in aqua regia only when alloyed with Pt. Oxidizes superficially by heating in O. Discovered by Wollaston in 1804.

Rhodium chloride (Rh_2Cl_6): Brownish red. Insoluble.

Rhodium cyanide $[\text{Rh}_2(\text{CN})_6]$: Carmine red.

Rhodium monoxide (RhO): Dark gray powder.

Rhodium trioxide (Rh_2O_3): Gray. Spongy. Insoluble.

Rhodium hydroxide $[\text{Rh}_2(\text{OH})_6]$: Brown. Gelatinous.

Rhodium sulphate $[\text{Rh}_2(\text{SO}_4)_3 + 12\text{H}_2\text{O}]$: Colorless, or yellowish crystals.

CII. OSMIUM.

Symbol, Os. Tetrad. Atomic weight, 198.494 (?). Specific gravity, 22.4. Melts, $2,500^\circ \text{C}$.

Infusible in oxyhydrogen flame. Somewhat volatile. Blue crystals. Hard. Oxidizes in air. Burns when heated.

Osmium chlorides: Bluish black OsCl_2 and orange OsCl_4 . Both soluble and made by heating Os in Cl. Os_2Cl_6 is only known in double salts.

Osmium and potassium cyanide $[\text{Os}(\text{CN})_2 \cdot 4\text{KCN} + 3\text{H}_2\text{O}]$: Colorless. Quadratic crystals.

Osmium monoxide (OsO): Dark insoluble powder.

Osmium trioxide (Os_2O_3): Black insoluble, and red scales.

Osmium dioxide (OsO_2): Red. Has metallic lustre.

Osmium tetroxide (osmic acid anhydride, OsO_4): Colorless prisms. One-per-cent solution used in hypodermic injection. Also as a microscopical reagent.

Osmamic acid: Has the formula $\text{Os}_2\text{O}_5\text{N}_2\text{H}_2$.

Osmous acid $[\text{OsO}_2(\text{OH})_2]$: Not known in free state; only in salts.

Osmium hydroxides: $\text{Os}(\text{OH})_2$ and $\text{Os}(\text{OH})_4$ as black powders.

Potassium osmite (K_2OsO_4): Unstable violet octohedra.

All osmium salts are very poisonous.

CIII. RUTHENIUM.

Symbol, Ru. Tetrad. Atomic weight, 103.56. Specific gravity, 12.26. Melts, $1,800^{\circ}\text{C}$.

Hard brittle crystals (from tin alloy as cubes) resembling Ir. Burns in O. Insoluble in acids, except very slightly in aqua regia. Dissolves as acid in melting KNO_3 , KClO_3 , or KOH.

Ruthenium dichloride (RuCl_2): Black and insoluble.

Ruthenium trichloride (Ru_2Cl_6): Deep yellow deliquescent crystals. Soluble in H_2O . When boiled it deposits $\text{Ru}_2(\text{OH})_6$.

Ruthenium tetrachloride (RuCl_4): Known only in double chlorides.

Ruthenium monoxide (RuO): Black and insoluble.

Ruthenium trioxide (Ru_2O_3): Does not blacken.

Ruthenium dioxide (RuO_2): Quadratic. Dark gray.

Ruthenium tetroxide (ruthenic acid anhydride, RuO_4): Yellow. Rhombic. Volatile at ordinary temperatures. Explodes with formation of black smoke when heated. Ozone odor.

Ruthenium hydroxides: $\text{Ru}_2(\text{OH})_6$ and $\text{Ru}(\text{OH})_4$.

Ruthenous acid: $\text{RuO}_2(\text{OH})_2$.

Perruthenic acid: $\text{RuO}_3(\text{OH})$.

Ruthenium sulphide: Ru_2S_3 .

RULES FOR THE ORTHOGRAPHY AND PRONUNCIATION OF CHEMICAL TERMS.

GENERAL PRINCIPLES OF PRONUNCIATION.

1. The pronunciation is as much in accord with the analogy of the English language as possible.
2. Derivatives retain as far as possible the accent and pronunciation of the root word.
3. Distinctly chemical compound words retain the accent and pronunciation of each portion.
4. Similarly sounding endings for dissimilar compounds are avoided (hence *-id*, *-ite*).

Accent.

In polysyllabic chemical words the accent is generally on the ante-penult; in words where the vowel of the penult is followed by two consonants, and in all words ending in *-ic* the accent is on the penult.

Prefixes.

All prefixes in strictly chemical words are regarded as parts of compound words, and retain their own pronunciation unchanged (as, *ă'ceto*, *ă'mīdo*-, *ă'zo*-, *nŷ'dro*-, *ī'so*-, *nī'tro*-, *nītro'so*-).

Elements.

In words ending in *-ium*, the vowel of the ante-penult is short if *i* (as *īrī'dium*), or *y* (as *didŷ'mium*), or if before two consonants (*că'lcium*), but long otherwise (as *tītă'nium*, *sělē'nium*, *chrō'mium*).

a'lŭminum	bī'smuth (biz)	că'lcium
a'ntimony	bō'ron	ca'rbon
a'rsenic	brō'min	cē'rium
bă'rium	că'dmium	cē'sium

* Adopted by the Section on Chemistry of the American Association for the Advancement of Science, 1891.

chlō'rin	magnē'sium	silicon
chrō'mium	(zhium)	silver
cō'balt	ma'nganese	sō'dium
colū'mbium	(eze)	strō'ntium
co'pper	me'rcury	(shium)
didŷ'mium	mōlŷ'bdenum	sŷ'lfur
e'rbiūm	nī'ckel	tā'ntalum
flū'orin	nī'trogen	tellŷ'rium
gā'llium	ō'smium	te'rbiūm
germā'nium	ō'xygen	thā'llium
glū'cinum	pallā'dium	thō'rium
gold	phō'sphorus	tin
hŷ'drogen	plā'tinum	tītā'nium
ī'ndium	potā'ssium	tŷ'ngsten
ī'ōdīn	rhō'dium	ūrā'nium
irī'dium	rubi'dium	vānā'dium
iron	ruthē'nium	ytte'rbiūm
lā'nthanum	samā'rium	ŷ'ttrium
lead	scā'ndium	zinc
lī'thium	sēlē'nium	zircō'nium

Also ämmō'nium, phosphō'nium, hā'logen, cyā'no-gen, ämŷ'dogen.

Note in the above list the spelling of the halogens, cesium and sulfur; f is used in the place of ph in all derivatives of sulfur (as sulfuric, sulfite, sulfo-, etc.).

Terminations in -ic.

The vowel of the penult in polysyllables is short (as cyā'nic, fūmā'ric, arsē'nic, siŷ'cic, iō'dic, būŷ'ric), except (1) u when not before two consonants (as mercū'ric, prŷ'ssic), and (2) when the penult ends in a vowel (as benzō'ic, olē'ic); in dissyllables it is long except before two consonants (as bō'ric, cī'tric).

Exceptions: acē'tic or acē'tic.

The termination -ic is used for metals only where there is a contrast with -ous (thus avoid aluminic, ammonic, etc.).

Terminations in -ous.

The accent follows the general rule (as plā'tinous, sŷ'lfurous, phō'sphorous, cōbā'ltous). Exceptions: acē'tous.

Terminations in -ate and -ite.

The accent follows the general rule (as *ă'cetâte*, *vă'nadâte*); in the following words the accent is thrown back (as *ă'bietâte*, *ă'lcoholâte*, *ă'cetonâte*, *ă'ntimonite*).

Terminations in -id (formerly -ide).

The final e is dropped in every case and the syllable pronounced -id (as *chlō'rīd*, *ī'odīd*, *hŷ'drīd*, *ō'xid*, *hydrō'xid*, *sū'lfīd*, *ă'mīd*, *ă'nīlīd*, *mūrē'xid*).

Terminations in -ane, -ene, -ine, and -one.

The vowel of these syllables is invariably long (as *mē'thāne*, *ē'thāne*, *na'phthalēne*, *a'nthracēne*, *prō'pine*, *quī'nōne*, *ă'cetōne*, *kē'tōne*).

The termination -ine is used only in the case of doubly unsaturated hydrocarbons, according to Hofmann's grouping.

Terminations in -in.

In names of chemical elements and compounds of this class, which includes all those formerly ending in -ine (except doubly unsaturated hydrocarbons) the final e is dropped, and the syllable pronounced -in (as *chlō'rīn*, *brō'mīn*, etc., *ă'mīn*, *ă'nīlīn*, *mo'rphīn*, *quī'nīn*, *vanī'llīn*, *alloxă'ntīn*, *absī'nthīn*, *emŷ'lsīn*, *că'ffeīn*, *cō'caīn*).

Terminations in -ol.

This termination, in the case of specific chemical compounds, is used *exclusively* for alcohols, and when so used is never followed by a final e. The last syllable is pronounced -ol (as *gly'cōl*, *phē'nōl*, *crēs'sōl*, *thy'mōl* (ti), *glŷ'cerōl*, *qui'nol*).

Exceptions: *ălcohōl*, *a'rgōl*.

Terminations in -ole.

This termination is always pronounced -ōle and its use is limited to compounds, which are not alcohols (as *ī'ndōle*).

Terminations in -yl.

No final e is used; the syllable is pronounced -ŷl (as ā'cetŷl, ā'mŷl, cē'rotŷl, ě'thŷl).

Terminations in -yde.

The y is long (as ā'ldehŷde).

Terminations in -meter.

The accent follows the general rule (hydrŏ'meter, barŏ'meter, lactŏ'meter).

Exception: Words of this class used in the metric system are regarded as compound words, and each portion retains its own accent (as cĕntime'ter, mī'llime'ter, kī'lome'ter).

Miscellaneous Words

which do not fall under the preceding rules.

Note the spelling: albumen, albuminous, albuminiferous, asbestos, gramme, radical.

Note the pronunciation: a'lkaline, a'lloy (n. and v.) a'llotropy, a'llotropism, ĭsomerism, pŏlymerism, apparā'tus (sing. and plu.), āqua regia, bary'ta, cĕn-tigrade, co'ncentrated, crystallĭn or crystalline, electrŏ'lysis, ĭter, mŏ'lecule, mŏlĕ'cular, nŏ'menclā'ture, olĕ'fiant, qua'ntivā'lence, vā'lence, ū'nivā'lent, bī'vā'lent, trī'vā'lent, qua'drivā'lent, tīt'rate.

A list of words whose use should be avoided in favor of the accompanying synonyms.

FOR	USE
beryllium	glucinum
niobium	columbium
thein	caffĕin
titer (n.)	strength or standard
titer (v.)	titrate
monovalent	univalent
divalent, etc.	bivalent, etc.
quantivalence	valence
sodic, calcic, zincic, nick- elic, etc, chlorid, etc.	sodium, calcium, zinc, nickel, etc., chlorid, etc. vid. terminations in -ic supra

FOR	USE
arsenetted hydrogen	arsin
antimonetted hydrogen	stibin
phosphoretted hydrogen	phosphin
sulfuretted hydrogen, etc.	hydrogen sulfid, etc.
alkylogens	alkylhaloids
benzol	benzene
toluol, etc.	toluene, etc.
pyrocatechin	catechol
resorcin	resorcinol
*hydroquinone (and hydrochinon)	quinol
orcein	orceinol
hydrophlorone	phlorol
phloroglucin	phloroglucol
quercite	quercitol
pinit	pinitol
glycerin	glycerol
erythrite, erythroglucin, eryglucin, erythromannite, phycite	erythrol
mannite	mannitol
dulcite	dulcitol
sorbite	sorbitol
furfurol	furfuraldehyde
fucosol	fucusaldehyde
anisol	methyl phenate
phenetol	ethyl phenate
anethol	methyl allyl-phenol

NOTE: It has been suggested that the words "qualitative" and "quantitative" could be advantageously replaced by "qualitive" and "quantitive," deriving the terms from the Latin adjectives instead of the

* Regarding this and the following words, cf. J. Chem. Soc. XLI., p. 248.

Fāte, fāt, fār, mōte, mōt, pīne, pīn, marīne, nōte, nōt, mōve, tūbe, tūb, rūle, my, ŷ=I.

' Primary accent; " secondary accent. N. B. The accent follows the vowel of the syllable upon which the stress falls, but does not indicate the division of the word into syllables.

nouns, as has been done in the case of "rotary" instead of "rotatory," "agriculturist" instead of "agriculturalist," etc. The Section regards this change as eminently desirable, but on account of the extended use of the words outside of chemistry, delays action until the opinions of those in allied branches have been obtained.

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LATIN AND GREEK NUMERALS

FOR

MEDICAL AND PHARMACEUTICAL STUDENTS.

Latin and Greek Numerals.

into many scientific terms Latin and Greek words enter as component parts. The beginner, who is not so acquainted with the classical languages, can only acquire such names and their meaning by memorizing, while he who knows the significance of the words or syllables, which form part of them, not only understands at once the meaning of the composite terms, but can more readily retain them in his memory. This, however, is but a very small part of the advantage derived from the study of the ancient languages, and the introduction of at least a short course of them, as a prerequisite to the study of the learned professions, is devoutly to be wished for. By a general agreement among the leading medical colleges of the United States, no student is to be admitted after 1883 who cannot pass an examination in easy Latin prose, at least. But many of those who have already entered are dedicated in this valuable auxiliary, and have no time left to them now to recover the lost ground. There, however, may much facilitate their studies by learning the meaning of, at least the most generally used foreign terms, employed in anatomy, histology, botany, natural philosophy, chemistry, etc.

For chemistry, perhaps more than for any other discipline of natural sciences, the knowledge of Latin and Greek numerals used in the nomenclature is almost indispensable. Fortunately, there are comparatively few needed and these may be readily learned. Some of the junior students of the Missouri Medical College, who were anxious to supply the deficiency in their preliminary studies, applied to Professor Dr. Carman for a brief table of the numerals, so as to aid them in better understanding and remembering the names of the different oxides, chlorides, etc., and the numeral terms used in crystallography, and he furnished them the following:

I. Cardinal Numbers.

ENGLISH.	LATIN.	GREEK.
1 One.	Unus, m. s.	Heis, m. s., hen (mone-).
2 Two.	duo, m. s.	Duo (doo) (di)
3 Three.	Tres, m. s.	Tres (tre)
4 Four.	Quatuor.	Tetras (tetra-)
5 Five.	Quinque.	Pent- (pente-)
6 Six.	Sex.	Hex.
7 Seven.	Septem.	Hept-.
8 Eight.	Octo.	Octo.
9 Nine.	Nove.	Nove.
10 Ten.	Decem.	Deca.
11 Eleven.	Undecim.	Endeca.
12 Twelve.	Duodecim.	Dodeca.
13 Thirteen.	Tredecim.	Trideca.
14 Fourteen.	Quatuordecim.	Tetradeca.
15 Fifteen.	Quindecim.	Pentadeca.
16 Sixteen.	Sedecim.	Hexadeca.
17 Seventeen.	Septendecim.	Heptadeca.
18 Eighteen.	Duodeviginti.	Dekadeca.
19 Nineteen.	Undeviginti.	Endekadeca.
20 Twenty.	Viginti.	Eikosa.
21 Twenty-one.	Viginti unus, etc.	Eikohene.
22 Thirty.	Triginta.	Triakonta.
23 Forty.	Quadraginta.	Tetrakonta.
24 Fifty.	Quingenta.	Pentakonta.
25 Sixty.	Sexaginta.	Hexakonta.
26 Seventy.	Septuaginta.	Heptakonta.
27 Eighty.	Octoginta.	Oktokonta.
28 Ninety.	Nonaginta.	Ennekonta.
29 Hundred.	Centum.	Hekaton (hekto-).
30 Two hundred.	Ducenti.	Dihekato.
31 Three hundred.	Trecenti.	Trihekato.
40 Four hundred.	Quadringenti.	Tetrasekato, etc.
500 Thousand.	Mille.	Chilio (chilo-).
1000 Ten thousand.	Decem milia.	Myrioi (myria-).

II. Ordinal Numbers.

ENGLISH.	LATIN.	GREEK.
1 First.	Primus.	Prontos.
2 Second.	Secundus.	Deterosus.
3 Third.	Tertius.	Triton.
4 Fourth.	Quartus.	Tetartus.
5 Fifth.	Quintus.	Pentopos.
6 Sixth.	Sextus.	Hexitos.
7 Seventh.	Septimus.	Hebdomos.
8 Eighth.	Octavus.	Oktos.
9 Ninth.	Nonus.	Ennekos.
10 Tenth.	Decimus.	Dekatos.
11 Eleventh.	Undecimus.	Eikotos.
12 Twelfth.	Duodecimus.	Triakotos.
13 Thirteenth.	Tredecimus.	Tetrakotos.
14 Fourteenth.	Quatuordecimus.	Pentakotos.
15 Fifteenth.	Quindecimus.	Hexakotos.
16 Sixteenth.	Sedecimus.	Hebdomakotos.
17 Seventeenth.	Septendecimus.	Oktokotos.
18 Eighteenth.	Octodecimus.	Ennekotos.
19 Nineteenth.	Novevicesimus.	Dekakotos.
20 Twentieth.	Vicesimus.	Triakakotos.
30 Thirtieth.	Tricesimus.	Tetrakakotos.
40 Four hundredth.	Quadringentesimus.	Pentakakotos.
500 Thousandth.	Millesimus.	Chiliakotos.
1000 Ten thousandth.	Decemvicesimus.	Myriakotos.

III. Denary Numerals.

ENGLISH.	LATIN.	GREEK.
A. Distributive.		
1 Single or one by one.	Singulis.	Monas.
2 Double or two by two.	Bial.	Dis or diobas.
3 Triple or three by three.	Terial or teral.	Trias or triobas.
4 Quadruple or four by four.	Quateral.	Tetrabias.
5, etc.	5, etc.	5, etc.

B. Multiplicative.

The English—fold is expressed in Latin by—*plic* (from *plico*, a fold), Greek—*plex*.

ENGLISH.	LATIN.	GREEK.
1 Simple (without fold).	Simplex.	Haplos.
2 Twofold.	Duplex.	Diplous.
3, etc.	3, etc.	3, etc.

C. Remunerative.

The English—times corresponds to the Greek—*his*.

ENGLISH.	LATIN.	GREEK.
1 Once.	Semel.	Hapax.
2 Twice.	Bis.	Dis.
3 Thrice.	Ter.	Trias or triakis.
4 Four times.	Quater.	Tetras.
5 Five times.	Quinquies.	Pentakis.
6 Six times.	Sexies.	Hexakis.
7 Seven times.	Septies.	Heptakis.
8 Eight times.	Octies.	Oktakis.
9, etc.	9, etc.	9, etc.

D. Miscellaneous Terms Relating to Quantity.

ENGLISH.	LATIN.	GREEK.
1 Half.	Semis (semi-).	Hemis.
10 One and one-half.	Sessqui.	Pnyx.
Many.	Mult.	Poly-
Large.	Magn.	Megalo-
Small.	Parva.	Mikro-

More—*basio*; di—*basio*; tri—*basio*; tetra—*basio*, used or acid to designate the number of hydrogen atoms replaceable by basic metals.

Tetra—*hedron*, crystal with four sides or facets; octa—*hedron*, crystal with eight sides or facets; dodeka—*hedron*, crystal with twelve sides or facets; eikosa—*hedron*, crystal with twenty-four sides or facets.

Triakis—*octo*—*hedron*, with three times eight facets; tetrakis—*hexa*—*hedron*, with four times six facets.

Hemi—*hedron*, crystal with one-half of the number of facets.

Duo—gramma	=	1-20 grammas.
Cent—gramma	=	1-100 grammas.
Milli—gramma	=	1-1000 grammas.
Deka—gramma	=	10 grammas.
Hekto—gramma	=	100 grammas.
Kilo—gramma	=	1000 grammas.
Myria—gramma	=	10000 grammas.

Mon—*oxide*, mono—*chloride*, having one atom of oxygen or chlorine in the molecule.

Di—*oxide* having two oxygen atoms. Tetra—*oxide* having four oxygen atoms.

Tetra—*chloride*, having four chlorine atoms. Penta—*sulphide*, having five atoms of sulphur in a molecule.

Quadri—*valent*, having four valences or atom-fixing powers.

Quadri—*oxy*, having four heads. Bi—*oxy*, having two heads.

Uni—*cellular*, consisting of a single cell.

Mult—*locular*, appearing in many places.

Diplo—*ocoon*, a micro—*ocoon* with double body, like an inverted figure 8.

Dicho—*tomos*, being split into two portions.

Poly—*eniphids*, compounds with many atoms of sulphur.

REPRINT FROM THE
MEYER BROTHERS DRUGGIST.
ST. LOUIS, MO

TABLE FOR CONVERTING CUSTOMARY INTO METRIC MEASURES AND WEIGHTS.

Published by MEYER BROTHERS DRUGGIST.

APOTHECARIES MEASURE.									
Cubic Min.	Cubic Fl. Oz.	Cubic Cm.	to	to	to	to	to	to	to
1 =	16.2305	0.47317	60.321	2.1135	0.20417	1 =	15.43130	0.517206	31.15074
2 =	32.461	0.94634	120.642	4.227	0.40834	2 =	30.8626	1.034412	62.30148
3 =	48.6916	1.41951	180.963	6.3405	0.61251	3 =	46.2939	1.551618	93.45222
4 =	64.9221	1.88268	241.284	8.454	0.81668	4 =	61.7252	2.068824	124.60296
5 =	81.1526	2.35585	301.605	10.5675	1.02085	5 =	77.1565	2.586030	155.75371
6 =	97.3831	2.82902	361.926	12.681	1.22502	6 =	92.5878	3.103236	186.90445
7 =	113.6136	3.30219	422.247	14.7944	1.42919	7 =	108.0191	3.620442	218.05520
8 =	129.8442	3.77536	482.568	16.9079	1.63336	8 =	123.4504	4.137648	249.20594
9 =	146.0747	4.24853	542.889	19.0214	1.83753	9 =	138.8817	4.654854	280.35668
APOTHECARIES WEIGHT.									
Gr.	to	to	to	to	to	to	to	to	to
1 =	0.035724	2.20462	0.070848	4.40964	0.141696	1 =	0.035724	2.20462	0.070848
2 =	0.071448	4.40924	0.143392	8.81928	0.283392	2 =	0.071448	4.40924	0.143392
3 =	0.107172	6.61386	0.215088	13.22892	0.425088	3 =	0.107172	6.61386	0.215088
4 =	0.142896	8.81848	0.286784	17.63856	0.566784	4 =	0.142896	8.81848	0.286784
5 =	0.17862	11.0231	0.35848	22.0482	0.70848	5 =	0.17862	11.0231	0.35848
6 =	0.214344	13.2277	0.430176	26.4578	0.860176	6 =	0.214344	13.2277	0.430176
7 =	0.250068	15.4323	0.501872	30.8674	1.011872	7 =	0.250068	15.4323	0.501872
8 =	0.285792	17.6369	0.573568	35.277	1.163568	8 =	0.285792	17.6369	0.573568
9 =	0.321516	19.8415	0.645264	39.6866	1.315264	9 =	0.321516	19.8415	0.645264
AVOIRDUPOIS WEIGHT.									
Gr.	to	to	to	to	to	to	to	to	to
1 =	0.035724	2.20462	0.070848	4.40964	0.141696	1 =	0.035724	2.20462	0.070848
2 =	0.071448	4.40924	0.143392	8.81928	0.283392	2 =	0.071448	4.40924	0.143392
3 =	0.107172	6.61386	0.215088	13.22892	0.425088	3 =	0.107172	6.61386	0.215088
4 =	0.142896	8.81848	0.286784	17.63856	0.566784	4 =	0.142896	8.81848	0.286784
5 =	0.17862	11.0231	0.35848	22.0482	0.70848	5 =	0.17862	11.0231	0.35848
6 =	0.214344	13.2277	0.430176	26.4578	0.860176	6 =	0.214344	13.2277	0.430176
7 =	0.250068	15.4323	0.501872	30.8674	1.011872	7 =	0.250068	15.4323	0.501872
8 =	0.285792	17.6369	0.573568	35.277	1.163568	8 =	0.285792	17.6369	0.573568
9 =	0.321516	19.8415	0.645264	39.6866	1.315264	9 =	0.321516	19.8415	0.645264

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